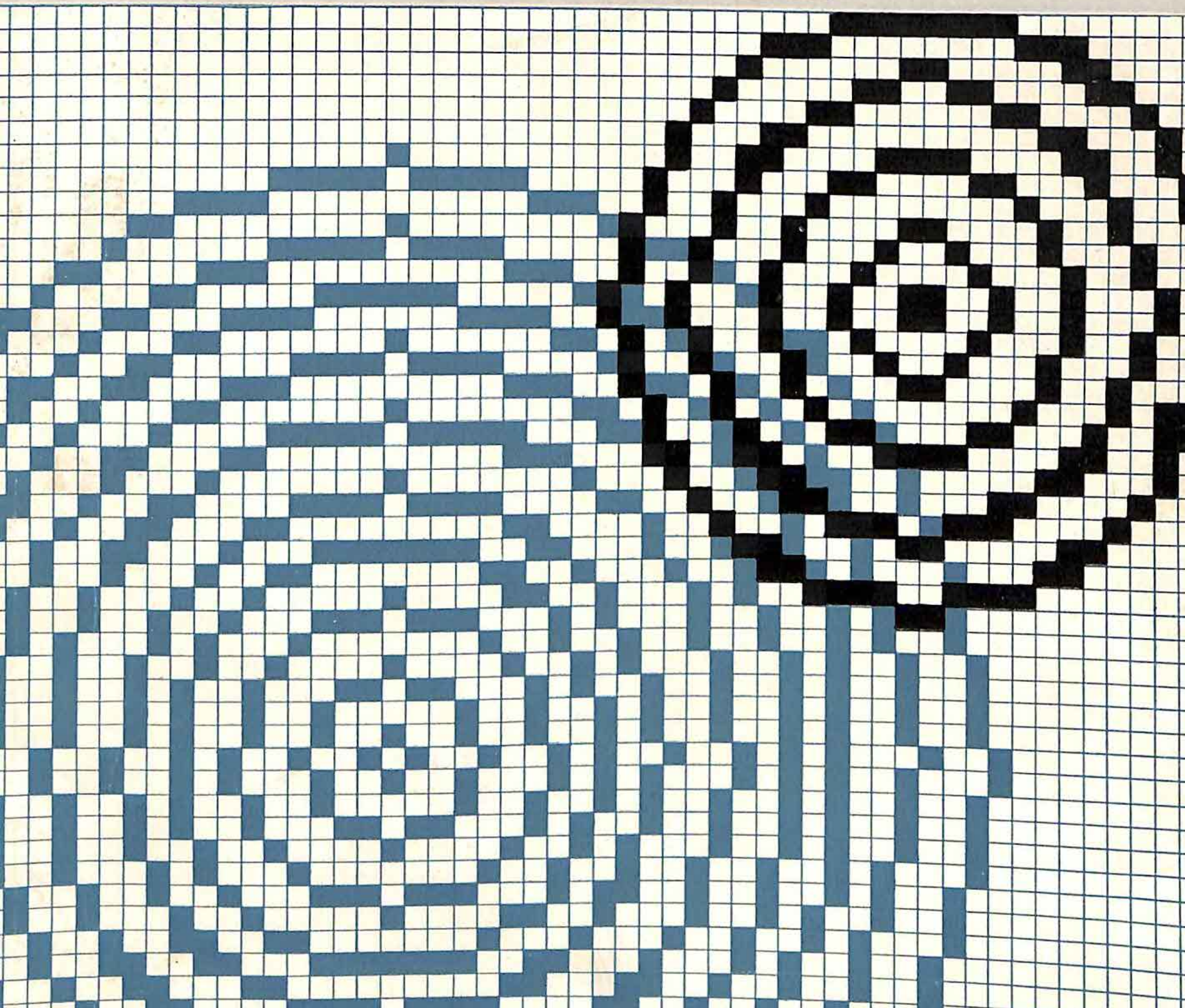
Understanding Science

CHEMISTRY

Class VIII



Reena Koshy • D.N. Varma • Vijayalakshmy Srivatsan

Understanding Science
CHEMISTRY

3

Reena Koshy
St. Xavier's School
Delhi

D N Varma
The Doon School
Dehradun

Editor
Vijayalakshmy Srivatsan
Vice Principal
P S Senior Secondary School
Madras

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Preface

The study of Chemistry is considered an endless memorization of innumerable chemical facts. It places very little emphasis on the most enjoyable part of the subject: experimentation, which really develops the manipulative and intellectual skills of students. The chemical principles grow out of careful observations of chemical changes and this attitude should be developed in students right from the beginning.

This series of three books is directed to achieve these aims. The students are encouraged to develop the habit of questioning and understanding, rather than being satisfied with blind acceptance of facts. At the end of this course students should be able to appreciate scientific methods.

The topics have been chosen according to the requirements of the syllabus laid down by the NCERT. We are certain that the approach will generate interest among the students for the study of the subject.

A large number of experiments and illustrative diagrams have been incorporated to help a better appreciation of the subject. The workbook portion at the end of each chapter will enable the students to assess their understanding of each topic.

November 1986

REENA KOSHY
D N VARMA
VIJAYALAKSHMY SRIVATSAN

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1

Carbon

Carbon occupies a unique position in the science of Chemistry. Carbon is the most important and common element of living matter and materials that support life. All kinds of foods, drugs, perfumes, dyes, plastics, synthetic fibres etc. contain carbon. Carbon is a constituent of numerous naturally occurring substances such as coal, mineral oils, carbonates, organic matter, etc. Carbon is present in the atmosphere in the form of carbon dioxide, as a small but important constituent. It would be correct to state that carbon is the central element in the structure of the biosphere—the part of the world in which life exists.

Millions of carbon compounds are already known and many more are being discovered or synthesised day after day. Due to this fact, a separate branch of Chemistry known as **Organic Chemistry**, is devoted to the study of carbon and its compounds.

Carbon has a unique ability of its atoms joining with one another to form very long chains or rings. Since each carbon atom of the chain can combine with atoms of other elements, the chain can form into a large molecule of a new compound. The number of carbon atoms in the chain can vary from one to infinity. Therefore, the number of compounds containing carbon are very many and abundant. Each molecule of octane (major constituent of petrol) contains 8 such carbon atoms; wax contains 25, while starch contains 100. Benzene contains 6 carbon atoms arranged in the form of a ring.

1.1 Forms of Carbon

Carbon exists in three different forms—diamond, graphite and amorphous carbon. While diamond and graphite are crystalline, amorphous carbon is non-crystalline as existing in charcoal, soot etc.

Diamond and graphite have different molecular structure and hence exhibit different physical properties.

Diamond and graphite are called **allotropes** of carbon. If an element can exist in two or more different physical forms but in the same physical state, the element is said to exhibit **allotropy.** Apart from carbon, phosphorus and sulphur also exhibit allotropy. Amorphous carbon is not an allotrope of carbon.

All forms of carbon show identical chemical properties. When burnt in presence of oxygen,

all of them give carbon dioxide
$C + O_2 \rightarrow CO_2$ + heat.

1.2 Diamond

Diamond is the purest form of carbon occurring naturally. Diamonds are colourless but often seen as coloured due to presence of slight traces of impurities. Diamonds were probably formed millions of years ago when carbon, dissolved in molten rock, crystallised under great pressure.

The carbon atoms are very strongly bonded to each other. Therefore, diamond is the hardest substance known. It does not conduct heat or electricity. Further, diamond does not normally react with any chemicals. But, when heated in presence of oxygen upto 900°C to 1000°C, diamond reacts with oxygen to produce carbon dioxide.
$C + O_2 \rightarrow CO_2$

Diamond has the property of causing internal reflection of rays of light and hence possesses a bright lustre. Due to the close packing of carbon atoms, diamond is the densest form of carbon, with a density of 3.5 g/cc.

Uses of Diamond

The two principal properties of diamond—lustre and hardness—make it very useful to mankind. The major uses are the following:
(1) As a precious gem in jewellery.
(2) As a drill head for mining and drilling rocks.

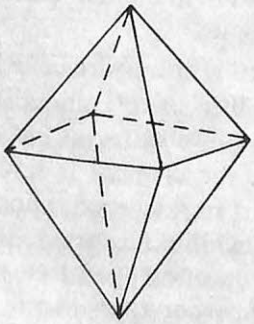


Fig. 1.1 A diamond crystal.

(3) As abrasive edge for cutting tools for metals and glasses.
(4) As a die for drawing thin wires, such as tungsten filament used in electric bulbs.

1.3 Graphite

Graphite also occurs naturally. The natural form of graphite contains certain impurities. A pure form of graphite can be prepared by heating coke (a substance obtained from coal) to a very high temperature in presence of inert gases in an electric furnace.

A graphite crystal is greyish black in colour. It has a soft and greasy texture. In a graphite crystal, carbon atoms are arranged in layers. The layers can slide over each other. This gives a lubricating property to graphite. Further, graphite is softer than diamond. Due to its softness and lubricating property, graphite can leave a mark on a surface.

Graphite has a metallic lustre but is opaque unlike diamond. It is a good conductor of electricity. The density of graphite is 2.2 g/cc. When heated to 700°C, it burns in oxygen to form carbon dioxide. Graphite can take part in a few chemical reactions.

Uses of Graphite

Based on the properties discussed above, graphite is most commonly used in the following applications.
(1) To manufacture pencil 'leads'
(2) As a lubricant for bearings.
(3) As an electrode.
(4) It is also used in nuclear reactors to regulate nuclear reactions.

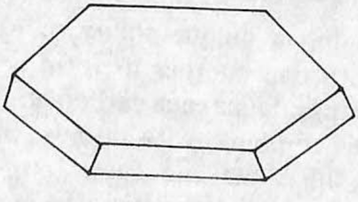


Fig. 1.2 A graphite crystal.

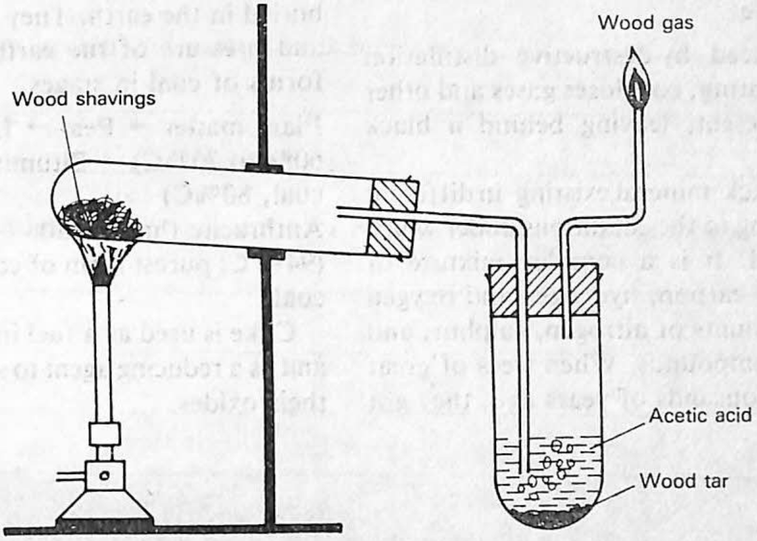


Fig. 1.3 Destructive distillation of wood in the laboratory.

1.4 Amorphous Carbon

The major amorphous forms of carbon are charcoal, lamp black and coke.

Charcoal

Charcoal is mainly obtained from three sources—wood, sugar and bone—by a process called **destructive distillation.** In this process, the source material is heated in the absence of air.

Wood charcoal is produced when wood is subjected to destructive distillation (Fig. 1.3). During this process, the volatile liquids and gases present in the wood escape, leaving behind black porous wood charcoal. Due to its porosity, this charcoal is able to absorb gases and is therefore used in gas masks and in water purification plants.

Animal or **bone charcoal** is obtained through destructive distillation of animal bones. This charcoal has the property of absorbing colouring matter. Hence it is used in industry as a decolourising agent. Common applications are manufacture of white sugar from brown sugar, white petroleum jelly and for purification of river water for town supply.

Sugar charcoal is a very pure form of carbon made by removing water from cane sugar or glucose using concentrated sulphuric acid, in absence of air.

$$\underset{\text{zinc oxide}}{ZnO} + C \xrightarrow{\text{heat}} \underset{\text{zinc}}{Zn} + \underset{\text{carbon monoxide}}{CO}$$

Sugar charcoal is mostly used as a reducing agent, to derive metals from their oxides.

$$\underset{\text{lead (II) oxide}}{PbO} + C \xrightarrow{\text{heat}} \underset{\text{lead}}{Pb} + \underset{\text{carbon monoxide}}{CO}$$

Lamp Black

Lamp black is formed when petrol, paraffin oil, turpentine etc. are burnt in limited supply of air. It is used as a colouring matter in preparation of Indian ink, black shoe polish, carbon paper, Printer's ink etc. Lamp black is also used in the manufacture of tyres to increase their durability. Soot is a coarse form of lamp black present in smoke coming out from chimneys of houses and factories.

Coal and Coke

Coke is produced by destructive distillation of coal. By heating, coal loses gases and other substances present, leaving behind a black residue—coke.

Coal is a black mineral existing in different forms according to the conditions under which it was formed. It is a complex mixture of compounds of carbon, hydrogen and oxygen with small amounts of nitrogen, sulphur, and phosphorus compounds. When trees of great forests died thousands of years ago, they got buried in the earth. They decayed under heat and pressure of the earth to form different forms of coal in stages.

Plant matter → Peat → Lignite (brown coal, 60% to 70%C) → Bituminous coal (domestic coal, 80%C)
Anthracite (hard coal)
(94% C; purest form of coal, called industrial coal)

Coke is used as a fuel in industrial furnaces and as a reducing agent to separate metals from their oxides.

EXERCISES

1 Why are compounds of no other element except carbon studied as a separate branch of Chemistry?

2. (i) What is allotropy?

(ii) Name three elements that exhibit allotropy.

(iii) Name the allotropes of carbon.

3. (i) Graphite and diamond exhibit different physical but the same chemical properties. Why?

(ii) Why is diamond such a hard substance? ______

______

______

(iii) Mention three important uses each of diamond and graphite. ______

______

______

______

______

4. Which property allows

(i) graphite to be used as an electrode? ______

(ii) diamond to be used in jewellery? ______

(iii) graphite to be used to make pencil leads? ______

(iv) diamond to be used in cutting tools? ______

(v) graphite to be used as a lubricant? ______

(vi) charcoal to be used in gas masks? ______

(vii) sugar charcoal and coke to be used to derive metals from their oxides? ______

______

5. (i) What do you mean by destructive distillation? ______

______

______

(ii) Name the main product obtained by destructive distillation of the following:

(a) Wood ______

(b) Animal bones ______

(c) Coal ______

6. Deposits of coal are found deep down in the earth. How do you think they are formed? ______

______

______

7. Complete and balance the following equations:

$ZnO + C \rightarrow$ ______________________________

$C + O_2 \rightarrow$ ______________________________

8. State true of false :
   (i) Carbon forms more compounds than any other element.
   (ii) Amorphous carbon contains small crystals of carbon.
   (iii) Diamond is denser than graphite.
   (iv) Sugar charcoal is made by destructive distillation of sugar.
   (v) Lamp black is used as a colouring matter in shoe polish.

9. Fill in the blanks:
   (i) The branch of chemistry dealing with the study of carbon and its compounds is ____________________ chemistry.
   (ii) ____________________ has the property of absorbing colouring matter. Hence it is used as a ____________________
   (iii) When petrol is burnt in a limited supply of air____________________ is formed.
   (iv) ____________________ is the hardest substance known.
   (v) ____________________ is used to regulate nuclear reactions in nuclear reactors.
   (vi) ____________________ coal is used as domestic fuel.

2

Compounds of Carbon—Oxides, Carbonates, Hydrocarbons

Carbon forms two oxides—carbon dioxide and carbon monoxide.

2.1 Carbon Dioxide (CO_2)

The atmosphere contains 0.03% by volume of carbon dioxide, as a natural constituent. Carbon dioxide is formed in the atmosphere due to the following natural processes:

(1) respiration of animals and plants;

(2) the burning of wood, coal, petrol, and such materials containing carbon;

(3) decay of vegetable matter;

(4) fermentation of alcohol, bread etc. where enzymes of yeast react with sugar and water, releasing carbon dioxide.

Carbon dioxide is prepared by (i) the action of acids on carbonates and bicarbonates:

$$\underset{\text{sodium bicarbonate}}{NaHCO_3} + \underset{\text{hydrochloric acid}}{HCl} \rightarrow \underset{\text{sodium chloride}}{NaCl} + H_2O + \underset{\text{carbon dioxide}}{CO_2}$$

$$\underset{\text{sodium carbonate}}{Na_2CO_3} + \underset{\text{sulphuric acid}}{H_2SO_4} \rightarrow \underset{\text{sodium sulphate}}{Na_2SO_4} + H_2O + CO_2$$

(ii) the action of heat on bicarbonates and carbonates:

$$\underset{\text{sodium bicarbonate}}{2NaHCO_3} \xrightarrow{\text{heat}} \underset{\text{sodium carbonate}}{Na_2CO_3} + H_2O + CO_2$$

$$\underset{\text{calcium carbonate}}{CaCO_3} \xrightarrow{\text{heat}} \underset{\text{calcium oxide}}{CaO} + CO_2$$

$$\underset{\text{copper carbonate}}{CuCO_3} \xrightarrow{\text{heat}} \underset{\text{copper oxide}}{CuO} + CO_2$$

Except the carbonates of sodium and potassium, all other carbonates decompose on heating to give carbon dioxide.

Laboratory Preparation of Carbon Dioxide

Carbon dioxide is prepared in the laboratory using marble or limestone ($CaCO_3$) and dilute hydrochloric acid or nitric acid. Dilute sulphuric acid is not often used since the calcium sulphate ($CaSO_4$) produced forms an insoluble coat around the carbonate, thus preventing further reaction.

$$\underset{\text{calcium carbonate}}{CaCO_3} + \underset{\text{nitric acid}}{2HNO_3} \rightarrow \underset{\text{calcium nitrate}}{Ca(NO_3)_2} + H_2O + \underset{\text{carbon dioxide}}{CO_2}$$

$$CaCO_3 + \underset{\text{hydrochloric acid}}{2HCl} \rightarrow \underset{\text{calcium chloride}}{CaCl_2} + H_2O + CO_2$$

$$CaCO_3 + \underset{\text{sulphuric acid}}{H_2SO_4} \rightarrow \underset{\text{calcium sulphate}}{CaSO_4} + H_2O + CO_2$$

In this experiment marble chips, and not powdered marble are used since the gas liberation would be too vigorous and uncontrollable if powder is used.

Activity 1: Set up the apparatus as shown in Fig. 2.1. Put marble chips very gently into the Woulfe's bottle or else it might crack. Pour dilute hydrochloric acid through the thistle funnel keeping the tip of the funnel inside the chips. Carbon dioxide starts bubbling out and gets collected in the gas jar by the upward displacement of air (carbon dioxide is heavier than air). To check whether the gas jar is full of carbon dioxide, introduce a burning splinter into the jar. If the flame is extinguished, it confirms that the jar is full of carbon dioxide. Collect several jars of carbon dioxide and keep them properly covered with glass lids.

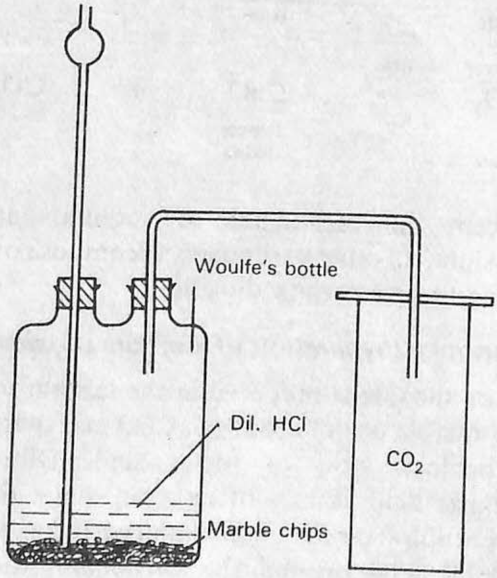


Fig. 2.1 Laboratory preparation of carbon dioxide.

Properties of Carbon Dioxide

Physical Properties

The carbon dioxide you have collected is a colourless gas. Try to smell the gas. You will find that it is also odourless.

Activity 2: Lower a burning candle on a gas jar spoon into a jar containing air. The candle would continue to burn. Now bring a jar containing carbon dioxide to the mouth of the first jar and tilt the jar as though pouring a liquid. The candle will be extinguished instantly. This shows that carbon dioxide 'flowed' downward into the first jar. This can only happen if carbon dioxide is heavier than air. In fact, carbon dioxide is 1½ times as dense as air. When compressed and cooled carbon dioxide readily forms a colourless liquid. If this liquid is allowed to evaporate rapidly, the cooling effect is sufficient to solidify (at – 78°C) the rest of the liquid to form solid carbon dioxide known as 'dry ice'.

A solution of carbon dioxide in water has a pleasant taste, as in soda water.

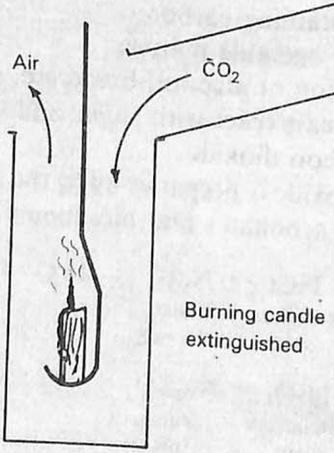


Fig. 2.2 Carbon dioxide is heavier than air.

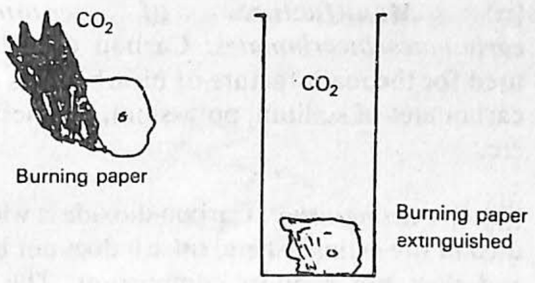


Fig. 2.3 Carbon dioxide does not support combustion.

Chemical Properties

Activity 3: Insert a burning paper into a jar of carbon dioxide. The flame will be extinguished instantly (Fig. 2.3). Carbon dioxide does not burn, nor does it support combustion. For this reason it is used widely as a fire extinguisher.

Activity 4: Burn a piece of magnesium ribbon, and introduce it into a jar of carbon dioxide. The ribbon continues to burn with aspluttering sound. When the magnesium is fully burned, a white powder of magnesium oxide will be left in the jar, in addition to black deposits of carbon. This happens because the magnesium flame is hot enough to decompose carbon dioxide to carbon and oxygen and the liberated oxygen helps the burning.

$$\underset{\text{magnesium}}{2Mg} + CO_2 \rightarrow \underset{\text{magnesium oxide}}{2MgO} + C$$

Activity 5: Invert a gas jar full of carbon dioxide in a trough containing water. You will observe that water rises in the jar (Fig. 2.4). This shows that carbon dioxide is soluble in water. In fact carbon dioxide is sparingly soluble in water. It dissolves in water to form a dilute solution of a weak acid known as carbonic acid.

$$CO_2 + H_2O \rightarrow \underset{\text{carbonic acid}}{H_2CO_3}$$

Carbon dioxide being an acidic oxide, reacts with solutions of alkalies to produce salt and water.

Activity 6: Pour a small amount of freshly prepared lime water into a jar of carbon dioxide. The lime water, which is colourless, turns milky white. If we allow the mixture to stand, the white precipitate would settle down.

$$\underset{\text{calcium hydroxide}}{Ca(OH)_2} + CO_2 \rightarrow \underset{\substack{\text{calcium carbonate}\\\text{(milky white)}}}{CaCO_3} + 2H_2O$$

If excess of carbon dioxide is passed through the milky solution, the mixture becomes clear again. This is due to the formation of soluble calcium bicarbonate.

$$CaCO_3 + H_2O + CO_2 \rightarrow \underset{\substack{\text{calcium}\\\text{bicarbonate}}}{Ca(HCO_3)_2}$$

A concentrated solution of calcium hydroxide is used for whitewashing the walls of buildings. The water evaporates leaving a white coat of moist slaked lime. The slaked lime absorbs carbon dioxide from the atmosphere and changes to milky white insoluble calcium carbonate giving a bright whiteness to the wall. As the days pass by, this calcium carbonate absorbs more carbon dioxide, along with water

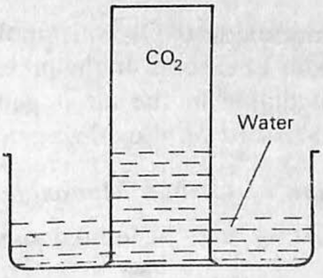


Fig. 2.4 Carbon dioxide is sparingly soluble in water.

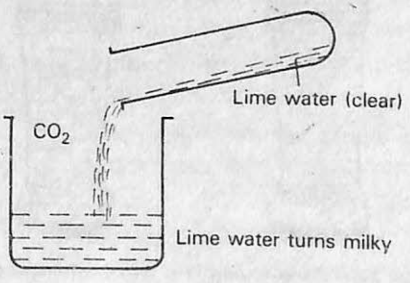


Fig. 2.5 Carbon dioxide turns lime water milky.

vapour from the atmosphere and gradually changes to soluble calcium bicarbonate. Calcium bicarbonate being soluble in water gets washed away during rains and finally it becomes time for another whitewash.

When carbon dioxide is passed over red hot coal, carbon monoxide is produced.

$$CO_2 + C \rightarrow \underset{\text{carbon monoxide}}{2CO}$$

Uses of Carbon Dioxide

(i) *Aerated drinks:* A large quantity of carbon dioxide dissolves in water when the gas is circulated under pressure. Part of the carbon dioxide reacts with water to form carbonic acid and the remaining stays as gas molecules mixed with the carbonic acid and water. This mixture, commonly known as **soda water** has a pleasant taste. Other aerated bottled drinks are also made in a similar manner by passing carbon dioxide through different flavoured solutions. When the lid of the aerated drink bottle is opened, the carbon dioxide escapes out of the solution causing effervescence (Fig. 2.6). It is interesting to note that effervescence seen while opening a beer bottle is not due to carbon dioxide applied under pressure before sealing the bottle. Here, it is due to the carbon dioxide produced within the sealed bottle due to continued fermentation of the beer.

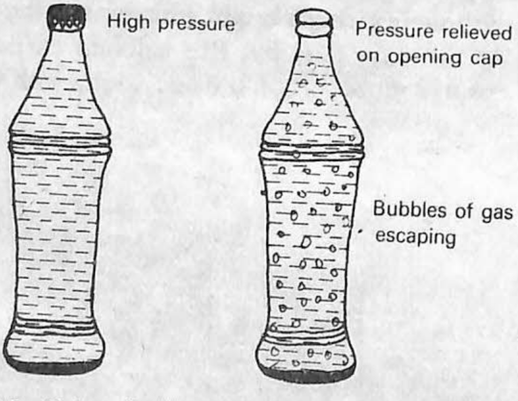


Fig. 2.6 Carbon dioxide is dissolved under pressure in soda water and other aerated cold drinks.

(ii) *Manufacture of common carbonates/bicarbonates:* Carbon dioxide is used for the manufacture of bicarbonates and carbonates of sodium, potassium, magnesium etc.

(iii) *Fire extinguisher:* Carbon dioxide is widely used in fire extinguishers, since it does not burn and does not support combustion. The gas being denser than air, displaces air and forms a blanket over the burning surface, and thus stops the fire.

(iv) *Refrigeration:* Solid carbon dioxide (dry ice) is used by ice-cream manufacturers as a refrigerant. Since the freezing point of carbon dioxide is $-78°C$, dry ice is a very effective refrigerant. Besides, dry ice is convenient to use as the solid ice on warming up directly changes to gaseous state. Artificial rain can be caused by air-dropping crushed dry ice into clouds to cool them enough to form rain drops.

(v) *Photosynthesis:* The most important use that nature puts carbon dioxide to is for the process of photosynthesis in plants. Plants absorb carbon dioxide from the atmosphere to manufacture their food by the process of photosynthesis. The carbon dioxide released by various sources to the atmosphere is used up by plants thus maintaining a proper balance of carbon dioxide in the atmosphere. Thus plants play an important role in reducing pollution level in the air.

2.2 Carbon Monoxide

Carbon monoxide (CO) is a highly unstable neutral oxide of carbon. In the presence of free oxygen available in the air it gets naturally converted to carbon dioxide.

Preparation of Carbon Monoxide

(a) By heating coke in limited supply of air.

$$2C + O_2 \xrightarrow{\text{heat}} 2CO$$

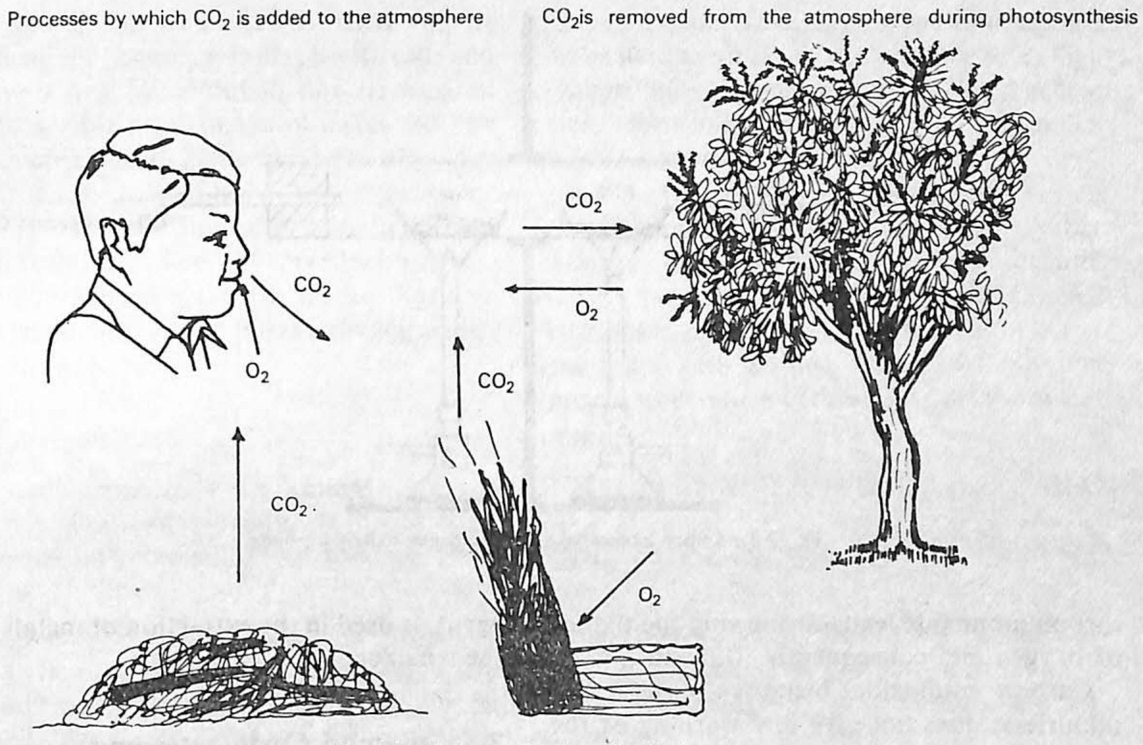

Processes by which CO2 is added to the atmosphere
CO2 is removed from the atmosphere during photosynthesis
CO2
O2
CO2
O2
CO2
CO2
O2

Fig. 2.7 Balance of carbon dioxide in nature

(b) By passing carbon dioxide over red hot charcoal.

$CO_2 + C \rightarrow 2CO$

(red hot)

Properties of Carbon Monoxide

Physical Properties

Carbon monoxide is a colourless, odourless, tasteless gas. Its density is slightly less than that of air. It is insoluble in water.

Chemical Properties

Carbon monoxide is a combustible gas. It burns with a blue flame to form carbon dioxide.

$$2CO + O_2 \rightarrow 2CO_2$$

Carbon monoxide is a good reducing agent. When it is passed over a heated metallic oxide, it removes oxygen from the metallic oxide to form carbon dioxide, leaving the metal behind.

$$CuO + CO \rightarrow Cu + CO_2$$

copper (II) oxide — copper

$$PbO + CO \rightarrow Pb + CO_2$$

lead (II) oxide — lead

Poisonous Nature of Carbon Monoxide

Carbon monoxide is a highly poisonous gas. When carbon monoxide mixed with air, in the ratio of 1 volume in 40,000 volumes of air is inhaled, it produces a violent headache; 1 volume in 800 volumes of air can cause death in half an hour, while 1 volume in 100 volumes of air can cause death in 2 minutes.

The haemoglobin present in blood combines with the oxygen we breathe to form oxyhaemoglobin. When carbon monoxide is inhaled, it combines with haemoglobin to produce carboxyhaemoglobin, which is formed at a much faster rate than oxyhaemoglobin. This compound is much more stable than oxyhaemoglobin and leaves very little haemoglobin for absorbing and circulating oxygen. This is how continued inhalation of

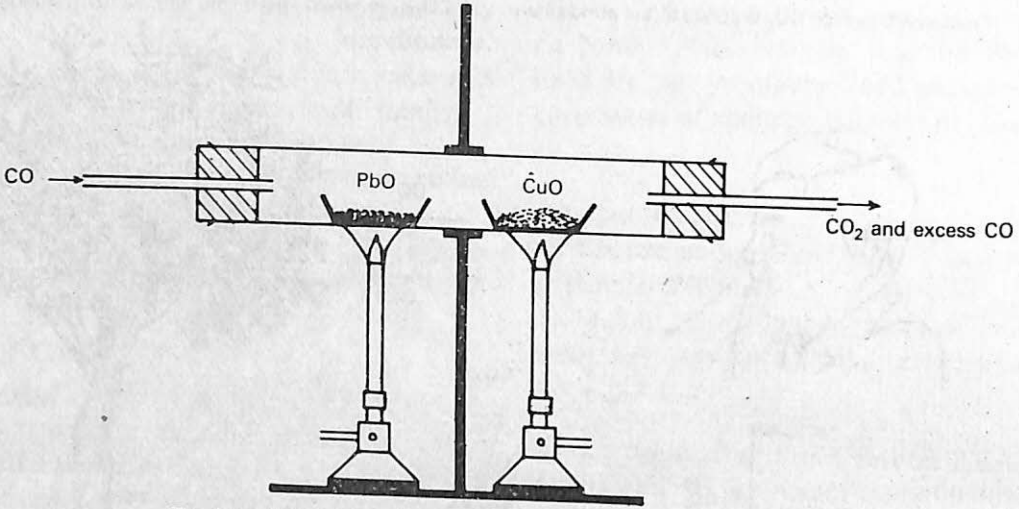


Fig. 2.8 Carbon monoxide can reduce oxides to form the metal

carbon monoxide leads to starving the tissues of oxygen and consequently to death.

Carbon monoxide being colourless and odourless, does not give any warning of the impending danger. At initial stages of breathing carbon monoxide, supply of excess of oxygen may help in reducing the ill effect. In severe cases of inhalation of carbon monoxide, blood transfusion may help.

Carbon monoxide is a danger hidden in many mines, factories and automobile garages. A very common source for carbon monoxide is a coal stove (*angithi*) burning in a closed room with limited supply of air. It is very important to provide adequate ventillation in the rooms to ensure oxygen supply in plenty, thereby minimising the formation of carbon monoxide.

Uses of Carbon Monoxide

(i) Carbon monoxide is an important constituent in gaseous fuel mixtures such as producer gas (mixture of carbon monoxide and nitrogen), water gas (mixture of carbon monoxide and hydrogen) and coal gas (mixture of carbon monoxide, methane and hydrogen).

(ii) Carbon monoxide, being a good reducing agent, is used in the extraction of metals from their oxides.

2.3 Common Carbonates and Bicarbonates

Sodium carbonate (Na_2CO_3) commonly known as **washing soda** is a white solid soluble in water. It is used in the manufacture of glass, soap, caustic soda and for softening hard water.

Calcium carbonate ($CaCO_3$) is a milky white solid insoluble in water. In chalk or limestone form, it is used for whitewashing buildings. Limestone, when mixed with linseed oil, forms a paste known as 'putty' which is for fixing window glasses. In marble chip form, calcium carbonate is used as a flooring material. Large pieces of marble are also used for the construction of exclusive buildings. Taj Mahal, for example, is made of marble.

Ammonium carbonate [$(NH_4)_2CO_3$] is used as a constituent in the preparation of smelling salts.

Sodium bicarbonate ($NaHCO_3$) is commonly known as **cooking soda**. It is a white solid soluble in water. It is used in the manufacture of baking powder, health salts and digestive tablets. When baking powder is mixed with dough for preparation of cakes etc. the bicarbonate reacts to give carbon dioxide which causes the cake to puff up. A similar action of release of carbon dioxide takes place in the case of Eno fruit salt, when it is mixed with water. Sodium bicarbonate present in the digestive tablets reacts with gastric juices reducing acidity in the stomach.

2.4 Hydrocarbons

These are compounds containing atoms of hydrogen and carbon only. Hydrocarbons existing naturally are numerous and important in our daily life.

Methane (CH_4) is the simplest form of hydrocarbon. This gaseous compound is the main constituent of a natural gas known as **marsh gas**. Methane is formed in marshy places due to the decomposition of vegetable matter under water. Methane is also formed in coal mines. Methane is a highly combustible gas and can be used as an excellent fuel. Due to its high combustibility the methane accumulated in coal mines when mixed with air sometimes causes disastrous mine explosions.

While 80 to 98% of natural gas obtained from oil wells is methane, the rest are other gaseous hydrocarbons, namely, ethane (C_2H_6), propane (C_3H_8), butane (C_4H_{10}). All these gases are colourless, insoluble in water and burn with a blue flame, without any smoke, to produce carbon dioxide and water vapour.

$$\underset{\text{methane}}{2CH_4} + 4O_2 \rightarrow 2CO_2 + 4H_2O$$

$$\underset{\text{butane}}{2C_4H_{10}} + 13O_2 \rightarrow 8CO_2 + 10\,H_2O$$

The cooking gas (L.P.G.) we use in our homes is a mixture of propane and butane.

Ethane is used in the manufacture of polythene.

Octane (C_8H_{18}) another important hydrocarbon, is the major constituent of petrol.

EXERCISES

1. State whether the following statements are true or false.
   (i) Dilute sulphuric acid is often used in the laboratory preparation of carbon dioxide.
   (ii) Carbon monoxide is a neutral oxide.
   (iii) Carbon dioxide is a combustible gas.
   (iv) Solid carbon dioxide is called 'dry ice'.
   (v) Carbon dioxide is a good reducing agent.
   (vi) Sodium bicarbonate is used in the manufacture of baking powder.
   (vii) Butane is the simplest hydrocarbon.
   (viii) Ethane is used to manufacture polythene.
   (ix) Sodium carbonate is commonly referred to as washing soda.
   (x) Sodium carbonate does not decompose on heating.

2. (a) (i) What are the processes by which carbon dioxide enters the atmosphere?______

____

____

(ii) How is the quantity of carbon dioxide in air maintained at 0.03% by volume?

(iii) Carbon dioxide does not support combustion. But a burning magnesium ribbon continues to burn in carbon dioxide. Explain.

(b) (i) Draw a labelled diagram to show the laboratory preparation of carbon dioxide.

(ii) Mention one other method by which carbon dioxide can be prepared.

(c) (i) How can you prove that carbon dioxide is heavier than air?

(ii) How can you prove that carbon dioxide is sparingly soluble in water? How can the solubility be increased?

____

____

____

____

____

(iii) Lime water turns milky when carbon dioxide is bubbled through it. The milkiness disappears on bubbling excess of carbon dioxide. Explain giving balanced equations. ______________________________

____

____

____

____

____

3. (i) Why is carbon dioxide used as a fire extinguisher? Can carbon monoxide be used instead? ______________________________

____

____

(ii) Why is it fatal to inhale a large quantity of carbon monoxide? ________________

____

____

____

____

4. Name the following.

(i) Two reducing agents. ______________________________

(ii) Two carbonates which do not liberate carbon dioxide on heating. ______

______

(iii) The process by which plants absorb carbon dioxide from the atmosphere and release oxygen. ______

(iv) The gas which burns with a blue flame to produce carbon dioxide on combining with oxygen. ______

(v) Two gaseous fuels in which carbon monoxide is one of the constituents. ______

______

(vi) The gas which sometimes causes mine explosions. ______

(vii) The gases present in L.P.G. (cooking gas). ______

______

(viii) A practical application of the property of carbon dioxide turning lime water milky.

______

______

(ix) The hydrocarbon which is the major constituent of petrol. ______

5. Complete and balance the following equations.

(i) $MgCO_3 \xrightarrow{heat}$ ______ + ______

(ii) $KHCO_3 + H_2SO_4 \rightarrow$ ______ + ______ + ______

(iii) $CO_2 + C \rightarrow$ ______

(iv) $PbO + CO \rightarrow$ ______

6. Fill in the blanks.

(i) ______ is an acidic oxide of carbon.

(ii) The mixture of carbon dioxide in water and carbonic acid is commonly called ______ .

(iii) ______ is used in the preparation of smelling salts.

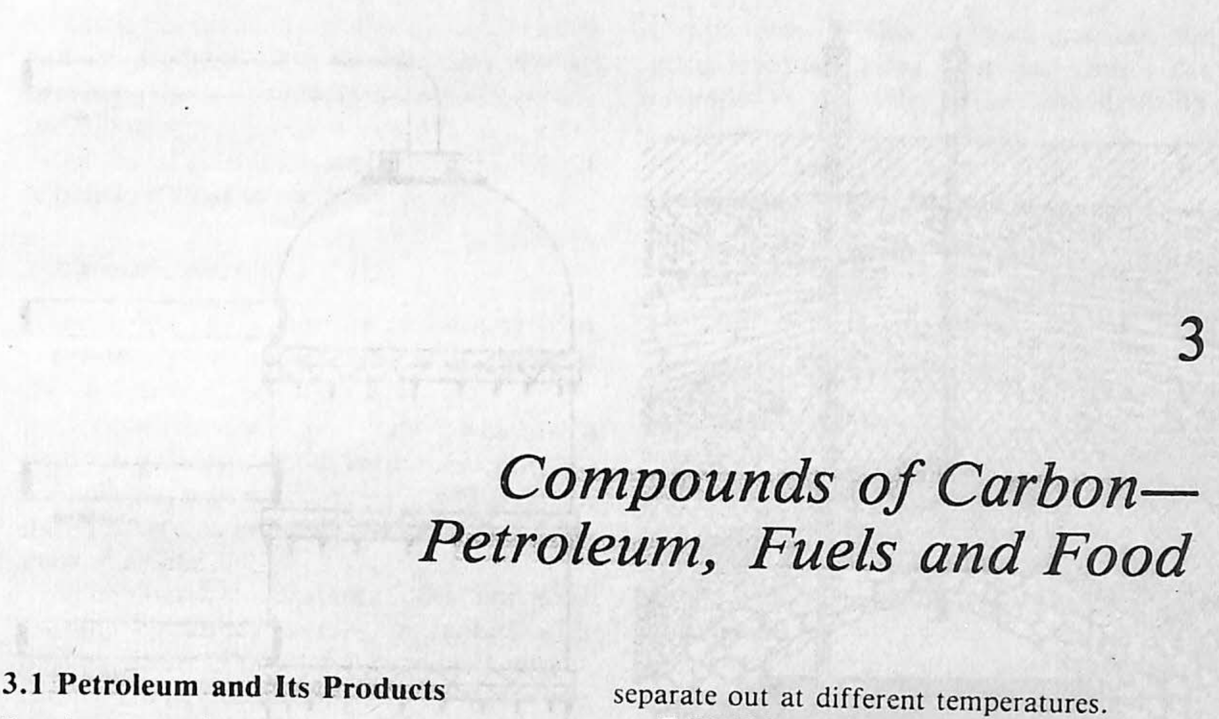
(iv) Digestive salts generally contain ______ .

(v) Solid carbon dioxide is known as ______ .

(vi) ______ is also known as marsh gas.

3

Compounds of Carbon—Petroleum, Fuels and Food



3.1 Petroleum and Its Products

Petroleum is a mixture of various hydrocarbons along with other organic compounds containing sulphur, nitrogen, oxygen, phosphorus etc. Petroleum, also referred to as **crude oil** is naturally occurring and dark in colour. The name 'petroleum' is derived from *petra* meaning rocks, and *oleum* meaning oil, signifying that it occurs under the earth's crust entrapped in rocks. The major producers of petroleum in the world are U.S.S.R., U.S.A., and Persian Gulf countries such as Saudi Arabia, Kuwait, Iran, Iraq, Egypt etc. India also produces petroleum, mainly from the oil wells of Assam, Gujarat and the Bombay offshore basin.

Crude oil is transported from the oil wells to the refineries through pipelines or tankers. At the refineries, crude oil is first washed with alkalies and sulphuric acid to remove the acidic and alkaline impurities. Thereafter it is refined by the process of **fractional distillation** to separate out different hydrocarbons present, taking advantage of their varying boiling points. The various components or **fractions** separate out at different temperatures.

TABLE 3.1 Main fractions in petroleum

Fraction	*Boiling point range*	*Uses*
Gases	Upto 40°C	Fuels.
Gasolene (naphtha)	40-200°C	Motor fuel (petrol), dry cleaning agent.
Kerosene	175-275°C	Fuel, lighting.
Diesel oil	200-300°C	Furnace oil, fuel for diesel engine.
Lubricating oil	Above 300°C	Lubrication.
Paraffin wax	Above 300°C	Candles, waxed paper, waterproof paper.
Asphalt, Pitch tar	Residue	Paving, roofing, fuel, paints and synthetic rubber.

In fractional distillation of the petroleum mixture, complete separation of the components is never possible. Hence redistillation of a fraction is done to derive the component in pure form. Petrol is produced through redistillation of naphtha.

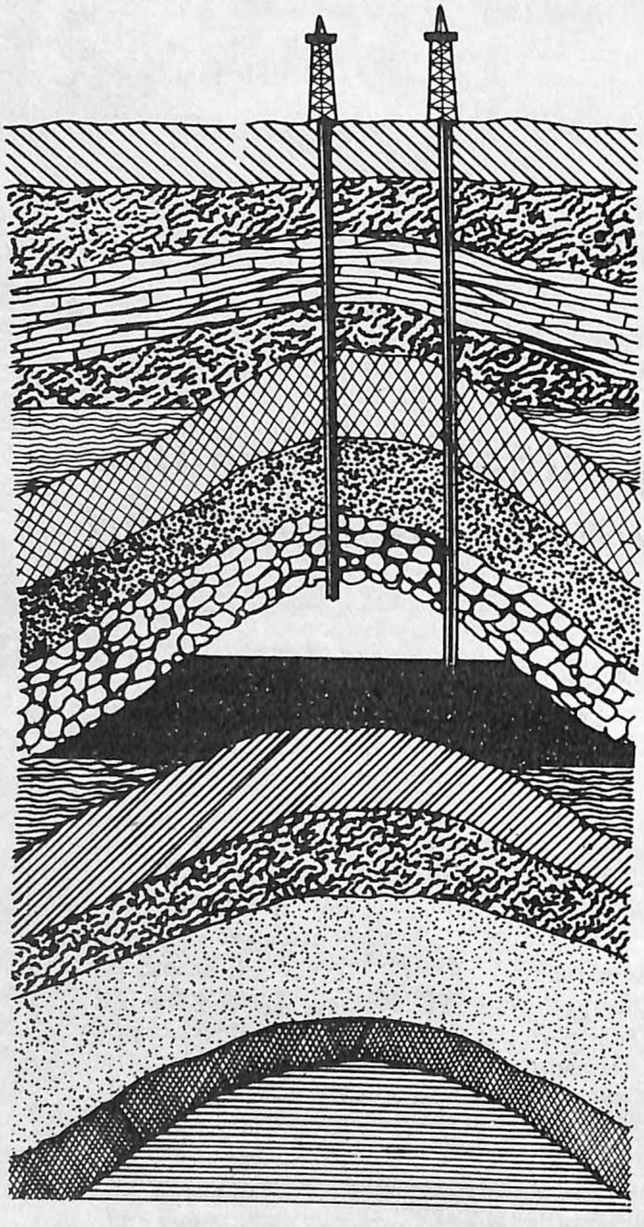


Fig. 3.1 An oil well

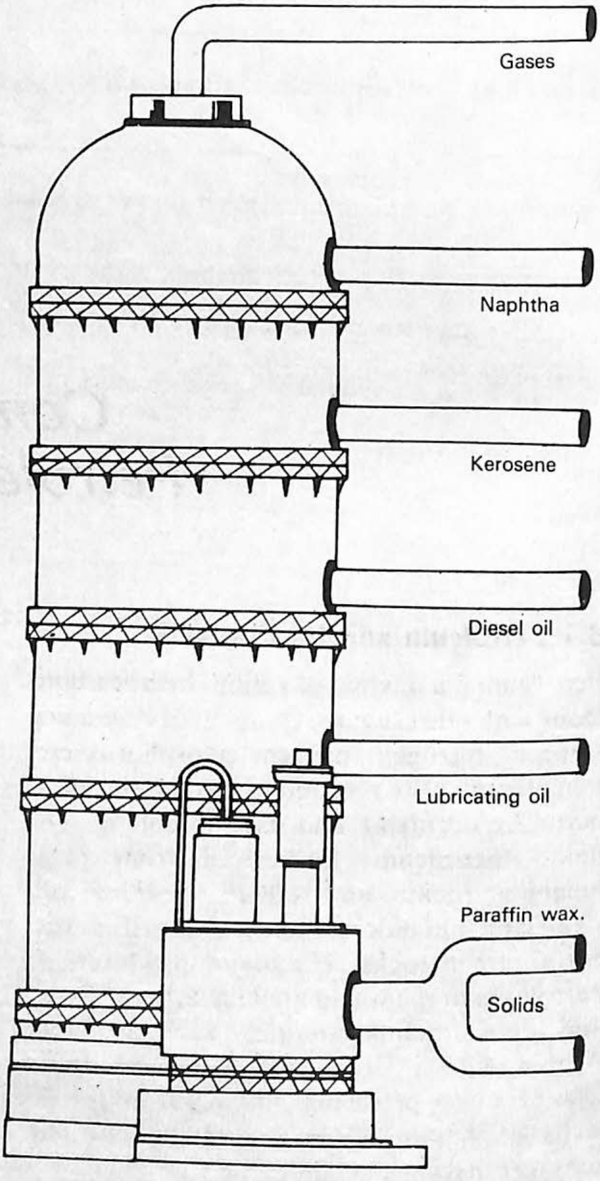


Fig. 3.2 Fractional distillation of petroleum.

Petroleum products are primarily used as fuels. Some of them are also used as lubricants for machinery. One of the products—wax—is extensively used for the manufacture of candles, polishes, waxed paper, water proofing etc. Some of the by-products are purified and used in the preparation of medicines, ointments, face creams and other cosmetics. Bitumen (asphalt) is used for surfacing of roads. From the fractions of petroleum, many more compounds are produced. These are insecticides, detergents, shampoos, plastics, synthetic rubber, perfumes, drugs, dyes, varnishes, photographic film, printer's ink, synthetic fibres etc. It is evident that petroleum and its products play an important role in providing us with comforts in daily life. In fact the economy of a nation depends to a great extent on its petroleum wealth. That is why it is often referred to as 'black gold'.

3.2 Combustion and Fuels

Combustion is the process of burning of a substance, generally in presence of oxygen, with the evolution of heat and light.

Combustible substances are those which burn readily e.g. paper, wood, kerosene, spirit, oil, cooking gas, wax etc. Substances which do not burn are termed **non-combustible,** e.g. stone, glass, clay etc.

A combustible substance does not start burning by simply coming in contact with oxygen. Paper or oil do not burn by themself. But when a lighted matchstick is brought close to them they start burning. This shows that a combustible substance starts burning only when heated to a certain minimum temperature. This lowest temperature at which a substance starts burning is known as its **ignition temperature** or **kindling temperature.** Burning ceases if the temperature of the combustible substance is lowered below its ignition temperature. This behaviour can be demonstrated by a simple experiment.

Activity 1: Light a candle. Surround the wick by a metal spring. You will find that the candle is extinguished. This happens because the spring conducts away heat and causes the temperature to fall below the ignition temperature of wax.

If combustion is associated with the evolution of sound energy, it is known as **explosion.**

Conditions necessary for combustion are

(1) A combustible substance,

(2) A supporter of combustion,

(3) Supply of heat to the combustible material to raise its temperature to the ignition temperature.

A fire can be extinguished by creating conditions contrary to combustion namely,

(1) Removing the combustible material.

(2) Cutting off the supply of air (supporter of combustion).

(3) Cooling the combustible material to bring down its temperature below the ignition temperature.

The most commonly used method of extinguishing a fire is to apply a suitable, non-combustible material which can simultaneously cut off the supply of air and cool the combustible substance. Water is an ideal fire extinguisher in most cases. However, water cannot be used for extinguishing fire in oil and petrol, since it is heavier, and oil and petrol can float on top of it and continue burning. It should also not be used to put off fires in electrical wiring as water conducts electricity and will cause electrical leakage.

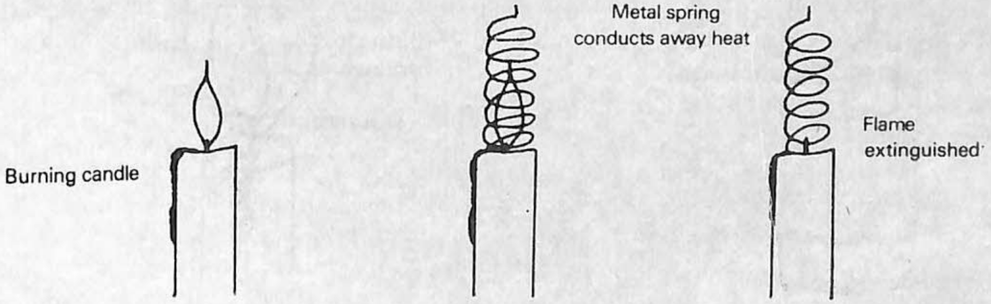


Fig. 3.3 A flame is extinguished if the temperature is brought down below the ignition temperature.

Fire Extinguishers

A **soda-acid** fire extinguisher consists of a metallic cylinder holding a solution of sodium bicarbonate. Sulphuric acid is kept in a separate glass bottle which can be broken by pressing in the knob. When we strike the knob against a hard surface, the glass bottle breaks. The consequent reaction between sodium bicarbonate and sulphuric acid produces water which acts as the cooling agent and carbon dioxide, which is a non-supporter of combustion. These are released under pressure towards the burning substance.

The **foam type** of fire extinguisher uses aluminium (III) sulphate and sodium bicarbonate along with a small quantity of *Turkey red oil.* When the constituents are made to react, a foam is produced which can effectively cut off the supply of air, to stop combustion.

Flame

Activity 2: Light a candle, and bunsen burner and observe the shape and colour of the flames. Do you notice any difference?

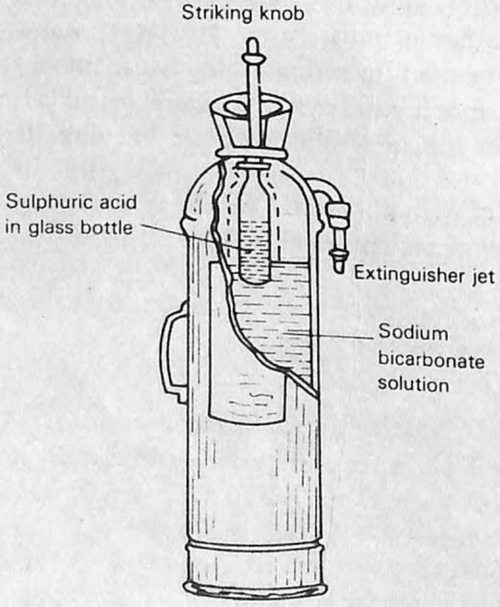


Fig. 3.4 A soda–acid fire extinguisher.

Now take a piece of magnesium ribbon and burn it. Compare with the burning of a candle, or a bunsen burner. What difference do you find?

You will notice that as magnesium burns it does not produce the conical region produced in case of a candle or a bunsen burner, which we call a **flame.** A flame is only produced when gases burn. Wax melts and vaporises to form a gas before burning. Magnesium does not vaporise and burns as a solid. In a spirit or kerosene lamp, the fuel vaporises before burning. Hence these lamps also produce flames.

Thus *a flame is a region in which gases combine with the evolution of heat and light.*

Structure of a Flame

A flame contains certain well defined regions within itself. Let us analyse the flame produced when a candle burns. Candle wax is a mixture of hydrocarbons. On application of heat the carbon and hydrogen atoms split up, each reacting with oxygen to produce carbon dioxide and water vapour. Formation of these two compounds is highly exothermic and hence a lot of heat is produced when wax burns. Different zones of the candle flame are marked in Fig. 3.5.

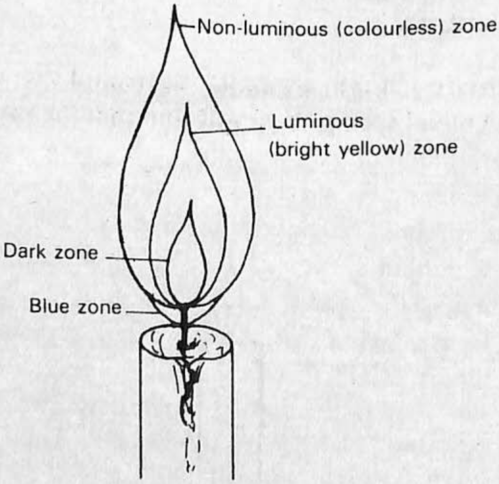


Fig. 3.5 A candle flame

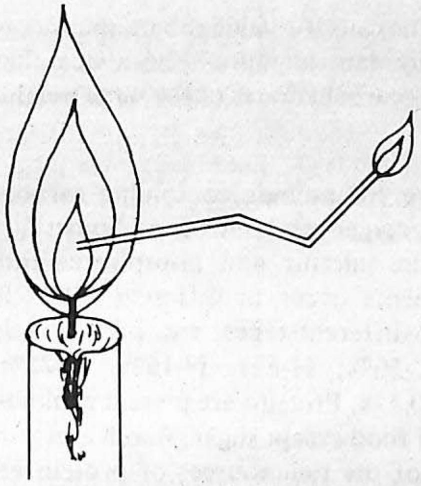


Fig. 3.6 The dark zone in a candle flame contains unburnt vapour.

(1) The central **dark zone** of the flame contains unburnt vapours of hydrocarbon. Burning does not take place here due to the temperature of the vapour not reaching its ignition level and non-availability of oxygen. If this vapour is drawn out by a glass tube it can be made to burn (Fig.3.6).

(ii) In the bright yellow zone, called the **luminous zone** there is incomplete combustion. The particles of unburnt carbon give yellow colour to the flame here.

(iii) In the outer zone, called the **non-luminous zone** combustion is complete and hence this is the hottest zone. It is only faintly visible.

(iv) In the lower part of the flame is a **blue zone** where the upward rush of air meets the combustible gas. This part receives plenty of air and burning is more complete than in the luminous zone.

3.3 Fuels

A fuel is any combustible substance used for producing heat energy by burning. Fuels are the most important source of energy for domestic and industrial uses.

Fuels can be solids, liquids or gases.
Solid fuels are coal, coke, charcoal, wood, wax.
Liquid fuels are kerosene, petrol, diesel, alcohol.
Gaseous fuels are methane, propane, butane, hydrogen, water gas, producer gas, acetylene etc.

A good fuel should be easily available, cheap, convenient to store and transport, controllable when burnt and should not produce undesirable gases and byproducts. Above all it should give a large amount of heat when burnt, i.e. it should be efficient.

Efficiency of a fuel is measured by its **calorific value.** Calorific value of a fuel is the amount of heat in calories produced by burning 1g of the fuel. Calorific values of a few common fuels are as follows:

Charcoal	800 cal/g
Wood	4000 cal/g
L.P.G.	10500 cal/g

3.4 Food

Food materials undergo combustion in our body, producing energy to maintain our body temperature and to enable us to do work. During winter season the body loses large amount of heat and hence the intake of food with high calorific value is necessary. The calorific value of foodstuffs is expressed in large calories (written as calorie with a capital C). In physics and chemistry we use the small calorie.

1 Calorie = 1000 calories.

The daily requirement of calories by a person depends mainly on the nature of work he does, climate and the age of the person. A person doing manual labour needs more calorie in-take than a clerk; we need more calories in winter than in summer; a young man needs more calories than an old man. A person of average weight (50-60 kg) needs about 1200 to 1400 Calories per day. However if he is doing hard manual work he would require 3000 to 3500

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Calories. It is interesting to note that a person lying quietly without movements spends energy at the rate of 1 Calorie per hour per kg of the body weight.

Chemistry of Food

The three major components of food are carbohydrates, fats and proteins. Their calorific values are

Carbohydrates	4.1 Cal/g
Fat	9.0 Cal/g
Protein	4.4 Cal/g

Carbohydrates

Carbohydrates are compounds of carbon, hydrogen and oxygen. They are generally represented by the formula $C_xH_{2x}O_x$ where x is an integral multiple of 6. Examples are glucose ($C_6H_{12}O_6$), sugar ($C_{12}H_{22}O_{11}$), starch or cellulose. Carbohydrates exist naturally in various food materials, such as mangoes, grapes, apples, beetroot, sugarcane, honey, wheat, rice and other cereals, potatoes etc.

The main function of carbohydrates is to supply energy. Carbohydrates of all categories get converted to glucose and get absorbed in the blood stream. This glucose reacts with the oxygen present in the blood stream to produce heat (energy), carbon dioxide and water. The unspent glucose gets stored in the liver and muscles to serve as an energy reserve.

Evidently, glucose in pure form is the fastest energy supplier, due to its capability to get into the blood stream directly. This is why athletes take glucose when exhausted.

Fats and Oils

Fats are also compounds of carbon, hydrogen and oxygen. Common forms of fat are butter, ghee, lard, fish oil, coconut oil, groundnut oil, mustard oil etc. They are classified into animal or vegetable fat depending on the source from which they are derived. Fats which exist in liquid form at room temperature are called oils.

The major function of fat is to produce energy. As the calorific value shows, the energy produced by fats is more than twice that generated by carbohydrates of the same weight.

Proteins

Proteins are compounds containing carbon, hydrogen, oxygen and nitrogen. Sometimes they contain sulphur and phosphorus too. These elements occur in different ratios in proteins of different types, e.g. milk protein contains C-50%, H-8%, N-16%, O-25%, S-0.5%, P-0.5%. Proteins are present in almost all forms of food except sugar, starch and pure fat. Some of the rich sources of protein are milk, fish, egg, meat, pulses and beans.

Proteins get converted to aminoacids by an enzyme in the body. In this form they are absorbed into various parts of the body where they recombine to form body tissues. The main function of proteins is to bring about body growth and repair of body tissues. For this reason children and convalescing patients require about 100g of protein daily.

Minerals

Minerals are salts of different elements such as iron, calcium, sodium, magnesium, manganese, copper, zinc, phosphorus, iodine etc. These are

TABLE 3.2 Calorific values of common food items

Food items	*Calorific value (Cal/100g)*
Pure carbohydrate	410
Pure fat	900
Pure protein	440
Bacon	570
Bread	260
Milk	70
Egg	150
Chocolate	550
Potatoes	90
Peas	100
Lettuce	18

required in small quantities for the proper functioning of the body. Minerals are available from milk, cheese, common salt and vegetables. Minerals of calcium and phosphorus help in building of bones and teeth. Presence of iron helps the blood in carrying oxygen and thereby in absorption of food for growth and energy.

Vitamins

Vitamins, which are complex compounds, along with minerals, regulate body processes and prevent diseases.

Our daily food must contain carbohydrates, fats, proteins, minerals, vitamins and water in appropriate quantities.

Calorific values of certain common food items are given in Table 3.2.

EXERCISES

1. (i) Under what condition can fractional distillation be used to separate components of a mixture? ______

   (ii) Name any four fractions of petroleum and write one use of each. ______

   (iii) Name the fraction of petroleum which is used as a domestic fuel. ______

2. (i) What are the conditions necessary for combustion to take place? ______

   (ii) What ways can be employed to extinguish a fire? ______

(iii) Which methods does a soda acid fire extinguisher use to put out fires?

(iv) Water cannot be used for extinguishing fire in oil and electrical wiring. Why?

3. (i) What is the reason for the existence of dark zone in a candle flame?

(ii) Is a flame produced when any material burns? Give relevant examples.

4. (i) What are the three major components of food? Give one example of each of them.

(ii) Growing children and convalescing patients should consume protein rich food. Why?

5. (i) In which unit is calorific value of food measured? ________________

(ii) How does this unit differ from the unit generally used in physics and chemistry? ____

____

(iii) In general, can a fat man survive much longer without food than a thin man? ____

____

____

____

____

6. Fill in the blanks.

(i) Refining of petroleum is done by the process of ________________

(ii) If combustion is associated with the release of sound energy it is said to be an ________________

(iii) ________________ is used with chemicals to produce foam in a foam type fire extinguisher.

(iv) The hottest zone of a flame is the ________________ zone.

(v) Calorific value of fat is ________________ times as much as that of an equal quantity of carbohydrates.

(vi) Combustion of food in our body produces energy to maintain our ________________ and to enable us to do ________________

7. Underline the correct terms.

(i) A fire extinguisher is a device to (produce, reduce, stop) combustion.

(ii) Burning ceases if the temperature of the combustible substance is lowered below its (melting point, kindling temperature, boiling point).

(iii) Incomplete combustion takes place in the (blue zone, dark zone, luminous zone) of a candle flame.

(iv) (Solid, Liquid, Gaseous) fuels have highest calorific value.

(v) Among all kinds of food (butter, glucose, milk) is the fastest energy supplier.

(vi) Proteins are not present in (pulses, fish, sugar).

8. Give an example each of a

(i) solid fuel ________________

(ii) liquid fuel ________________

(iii) a source of protein. ________________

4

Structure of the Atom

You have read about Dalton's Atomic Theory. Dalton considered all matter to be composed of atoms. He believed that atoms were the smallest indivisible particles of an element. One of the important discoveries of the twentieth century, however, was that the atom itself has a sub-structure of smaller particles.

4.1. Electrons, Protons and Neutrons

Atoms are extremely small. Their average diameter is only 10^{-8}cm. One gram of a substance contains several million atoms.

Thomson conducted experiments with electrodes placed in gases under very low pressure (the apparatus was known as **cathode ray tube**). Through these experiments he was able to prove that all atoms contained **electrons,** which were negatively charged and had a mass of only 1/1835 of that of an atom of hydrogen. The charge on an electron is taken to be unit negative charge.

Rutherford in 1920 discovered **protons** which has a unit positive charge. Their mass was the same as the mass of a hydrogen atom. In 1932 Chadwick discovered a neutral particle with the same mass as that of a proton. He called it **neutron.**

Every atom contains electrons, protons and neutrons (except hydrogen which has no neutron). The properties of the three fundamental particles are summarised in Table 4.1.

TABLE 4.1 Properties of sub-atomic particles

Particle	*Charge*	*Mass*	*Represented by*
Electron	−1	1/1835 of hydrogen atom	e
Proton	+1	same as hydrogen atom	p
Neutron	Neutral	same as hydrogen atom	n

4.2 Atomic Model

Several models of the atom were proposed before Rutherford, in 1911, conducted his famous gold leaf experiment. He showed that most of the space in the atom was empty. The mass and positive charge of the atom were concentrated in a small region.

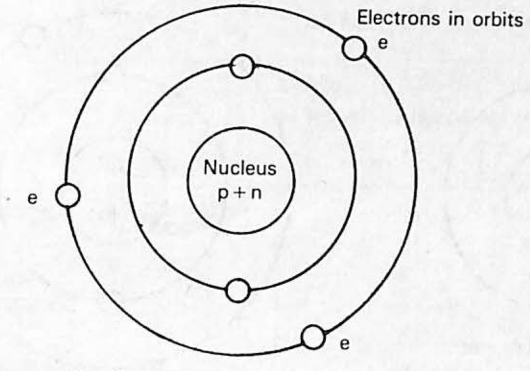


Fig. 4.1 Bohr's model of an atom.

Based on Rutherford's experiments Niels Bohr, in 1913, suggested that the atom had a **nucleus** in the centre, consisting of protons and neutrons closely packed together. Electrons move around the nucleus in definite orbits (Fig. 4.1). The electrons are kept from flying off from the orbits by the attraction existing between the negatively charged electrons and the positively charged protons in the nucleus. The nucleus occupies a very small volume.

This model is very similar to the solar system with the sun forming the nucleus, and the planets the electrons.

4.3 Atomic Number and Mass Number

As the atom as a whole is neutral, the number of protons in the nucleus of an atom is the same as the number of electrons in the orbits. This number is called the **atomic number** of the atom. It determines the chemical properties of the atom. The atomic number is written at the left bottom corner of the chemical symbol of an element.

${}_{6}C$ represents an atom of carbon with atomic number = 6

${}_{20}Ca$ represents an atom of calcium with atomic number = 20

${}_{92}U$ represents an atom of uranium with atomic number = 92

We have seen that an electron weighs only 1/1835 of the mass of a proton. From a chemical viewpoint this is considered to be negligible or zero. Imagine a sugar candy ball with dust particles sticking on it. We would not notice any change in the mass of the candy ball with or without dust particles. In a similar manner, the presence of electrons around a nucleus cannot show any noticeable change in mass.

Thus, the mass of an atom is due to the protons and neutrons only. The sum of the number of protons and neutrons represents the **mass number** of the atom. The mass number is written at the left top corner of the chemical symbol.

4.4 Relative Atomic Mass

Atomic mass is the mass of an atom of the element. Since the atom is too small, its mass cannot be measured on a balance. However, physicists have been able to find out the weights of atoms of different elements by indirect methods. An atom of hydrogen weighs 1.67×10^{-24}g and an atom of oxygen weighs 26.5×10^{-24}g. Thus if we were to write the actual weights of atoms of elements in grams, it would mean working with very small numerical factors.

Just as we have units for length and mass, e.g. a metre and a kilogram respectively, for comparing the lengths and masses of substances, we could take the mass of some light particle as a unit to compare the mass of atoms.

Scientists have accepted ${}^{12}_{6}C$ as—the standard atom and one-twelfth (1/12) of the mass of ${}^{12}_{6}C$ is accepted as one **atomic mass unit.**

Relative atomic mass *is the mass of an atom of an element when compared to the standard atomic mass unit.* Since this is a relative measurement, relative atomic mass is a number without a unit.

${}^{12}_{6}C$ represents a carbon atom with mass number = 12 and atomic number = 6

${}^{35}_{17}Cl$ represents a chlorine atom with mass number = 35 and atomic number = 17

${}^{235}_{92}U$ represents a uranium atom with mass number = 235 and atomic number = 92

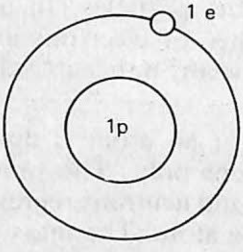


Fig. 4.2 Hydrogen atom.

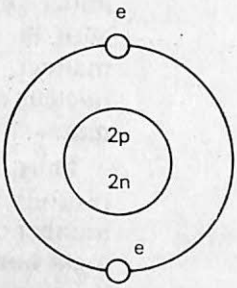


Fig. 4.3 Helium atom.

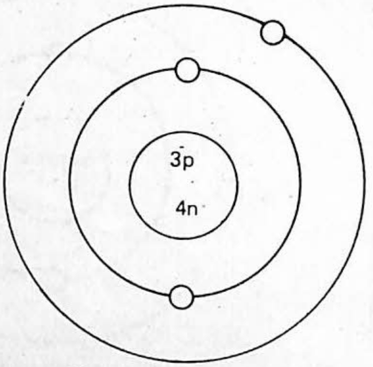


Fig. 4.4 Lithium atom.

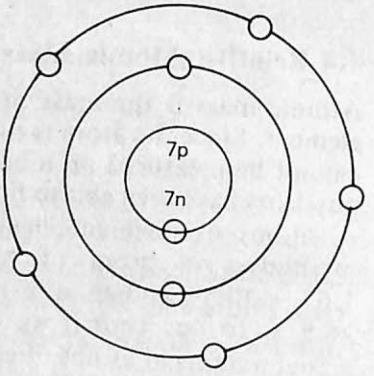


Fig. 4.5 Nitrogen atom

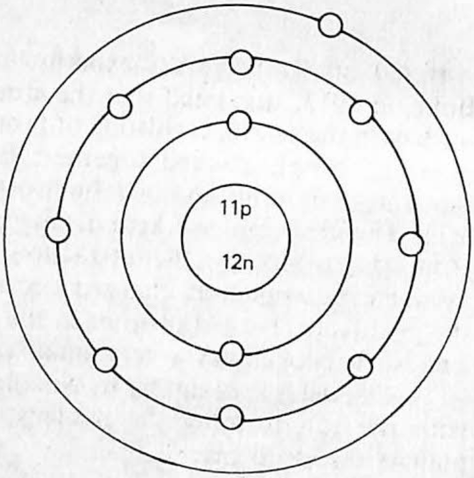


Fig. 4.6 Sodium atom

4.5 Electrons in an Atom

Electrons are arranged in different **orbits** or **shells (also known as energy levels)**. The shells are numbered 1, 2, 3 etc. outwards from the nucleus. Each orbit can have a certain maximum number of electrons. The first (innermost) orbit can hold up to 2 electrons, the second orbit upto 8 electrons, the third orbit upto 18 electrons, the fourth orbit upto 32 electrons and so on. Electrons in the different orbits possess different levels of energy. Electrons of the innermost orbit have least energy while those in the outermost orbit have maximum energy.

Once an orbit is full the electrons occupy the next orbit until that is full, and so on. However, the outermost orbit cannot have more than 8 electrons.

Hydrogen has only one proton and one electron. The electron is present in the first orbit.

Helium has two protons, two neutrons and two electrons. The two electrons occupy the first orbit, filling it up completely.

Lithium has three protons, four neutrons and three electrons. Two of the electrons occupy the first orbit and the third electron goes to the second orbit. We write the electronic configuration of lithium as
Lithium 2, 1

Nitrogen has seven electrons. Two of these will go in the first orbit and five in the second orbit. The electronic configuration of nitrogen is
Nitrogen 2, 5

Sodium has eleven electrons. Two of these go in the first orbit, eight in the second orbit and

TABLE 4.2 The first twenty elements

Name	*Symbol*	*Atomic number*	*No. of neutrons*	*Mass number*	*Relative atomic mass*	*Electronic configuration*
Hydrogen	H	1	0	1	1.008	1
Helium	He	2	2	4	4.003	2 First shell complete
Lithium	Li	3	4	7	6.941	2, 1 Second shell starts
Beryllium	Be	4	5	9	9.012	2, 2
Boron	B	5	6	11	10.81	2, 3
Carbon	C	6	6	12	12.01	2, 4
Nitrogen	N	7	7	14	14.01	2, 5
Oxygen	O	8	8	16	16.00	2, 6
Fluorine	F	9	10	19	19.00	2, 7
Neon	Ne	10	10	20	20.18	2, 8 Second shell complete
Sodium	Na	11	12	23	22.99	2, 8, 1 Third shell starts
Magnesium	Mg	12	12	24	24.31	2, 8, 2
Aluminium	Al	13	14	27	26.98	2, 8, 3
Silicon	Si	14	14	28	28.09	2, 8, 4
Phosphorus	P	15	16	31	30.97	2, 8, 5
Sulphur	S	16	16	32	32.06	2, 8, 6
Chlorine	Cl	17	18	35	35.45	2, 8, 7
Argon	Ar	18	22	40	39.95	2, 8, 8 Third shell has 8 electrons
Potassium	K	19	20	39	39.10	2, 8, 8, 1 Fourth shell starts
Calcium	Ca	20	20	40	40.08	2, 8, 8, 2

one in the third orbit. The electronic configuration is
Sodium 2, 8, 1
Notice that as the number of shells increase, the size of the atom increases.

4.6 Isotopes

All atoms of one element have the same number of protons and electrons and hence the same atomic number. However, they may have different number of neutrons, and therefore different masses. In such cases the chemical properties do not change as these depend on the electrons in the atom.

Atoms of the same element, which have the same chemical properties, but different mass numbers are called **isotopes.**

Hydrogen has three isotopes containing 0, 1, and 2 neutrons respectively.

Hydrogen $^{1}_{1}H$
Deuterium $^{2}_{1}H$
(heavy hydrogen)
Tritium $^{3}_{1}H$

The atomic mass of natural hydrogen is 1.008 as it contains a small amount of deuterium and tritium.

Chlorine has two isotopes containing 18 and 20 neutrons with mass numbers 35 and 37 respectively.

$^{35}_{17}Cl$
$^{37}_{17}Cl$

In ordinary chlorine there are about three times as many $^{35}_{17}Cl$ atoms as $^{37}_{17}Cl$ atoms. The

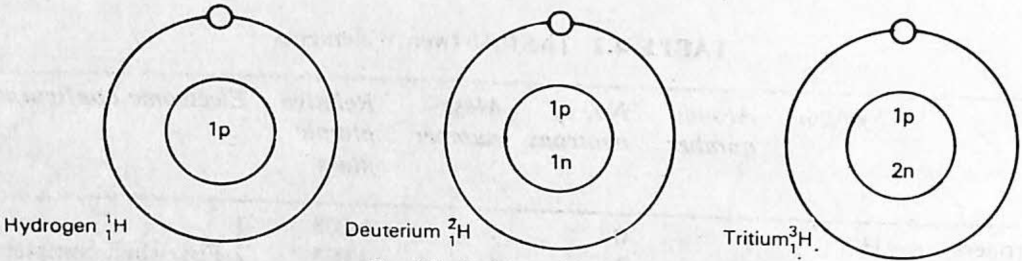


Fig. 4.7 Isotopes of hydrogen.

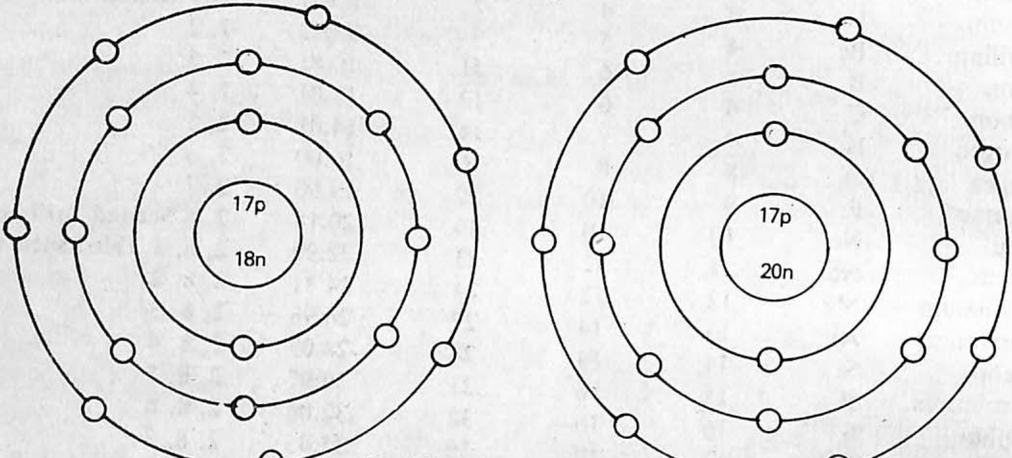


Fig. 4.8 Isotopes of chlorine

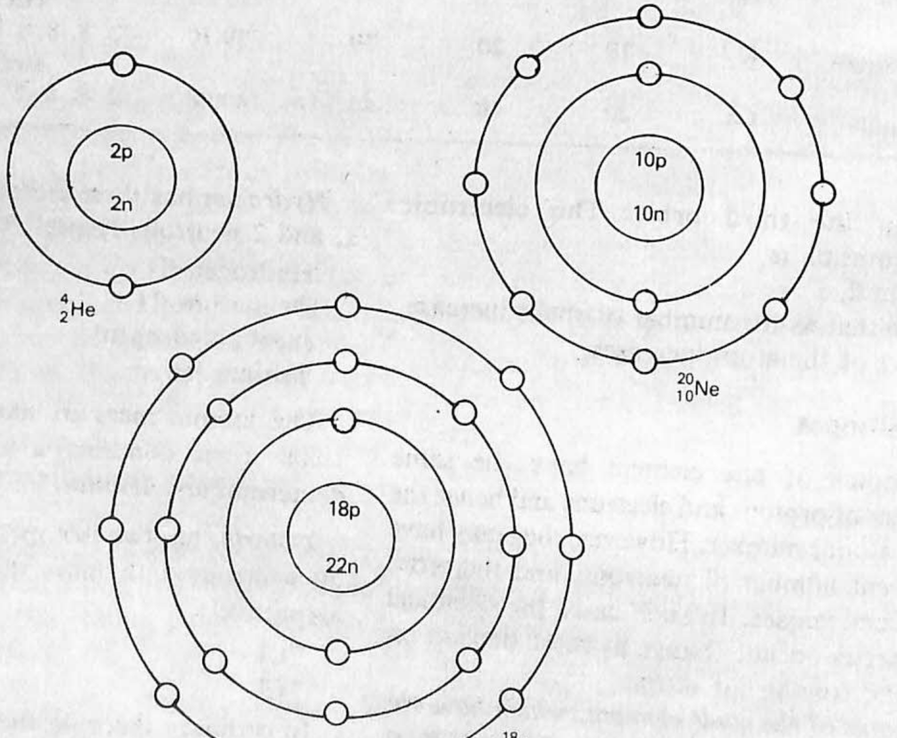


Fig. 4.9 Electronic configurations of stable, unreactive elements.

average mass number of chlorine is therefore 35.5.

Some other common elements with isotopes are carbon ($^{12}_{6}\text{C}$ and $^{14}_{6}\text{C}$) and oxygen ($^{16}_{8}\text{O}$, $^{17}_{8}\text{O}$ and $^{18}_{8}\text{O}$)

The existence of isotopes explains why several elements have fractional mass numbers.

4.7 Electrons and Chemical Reactivity

Atoms which have either a full outer electron orbit or have 8 electrons in their outermost orbit are stable and very unreactive. Examples are helium (full outermost orbit), neon (full outermost orbit) and argon (8 electrons in outermost orbit) (Fig. 4.9).

These are known as **inert** or **noble gases.**

In any chemical reaction, atoms gain, lose or share electrons in an attempt to achieve maximum stability by having an electronic configuration of a noble gas.

When sodium (2, 8, 1) combines with chlorine (2, 8, 7), the sodium atom loses one electron and the chlorine atom gains it. Thus the electronic configurations change to

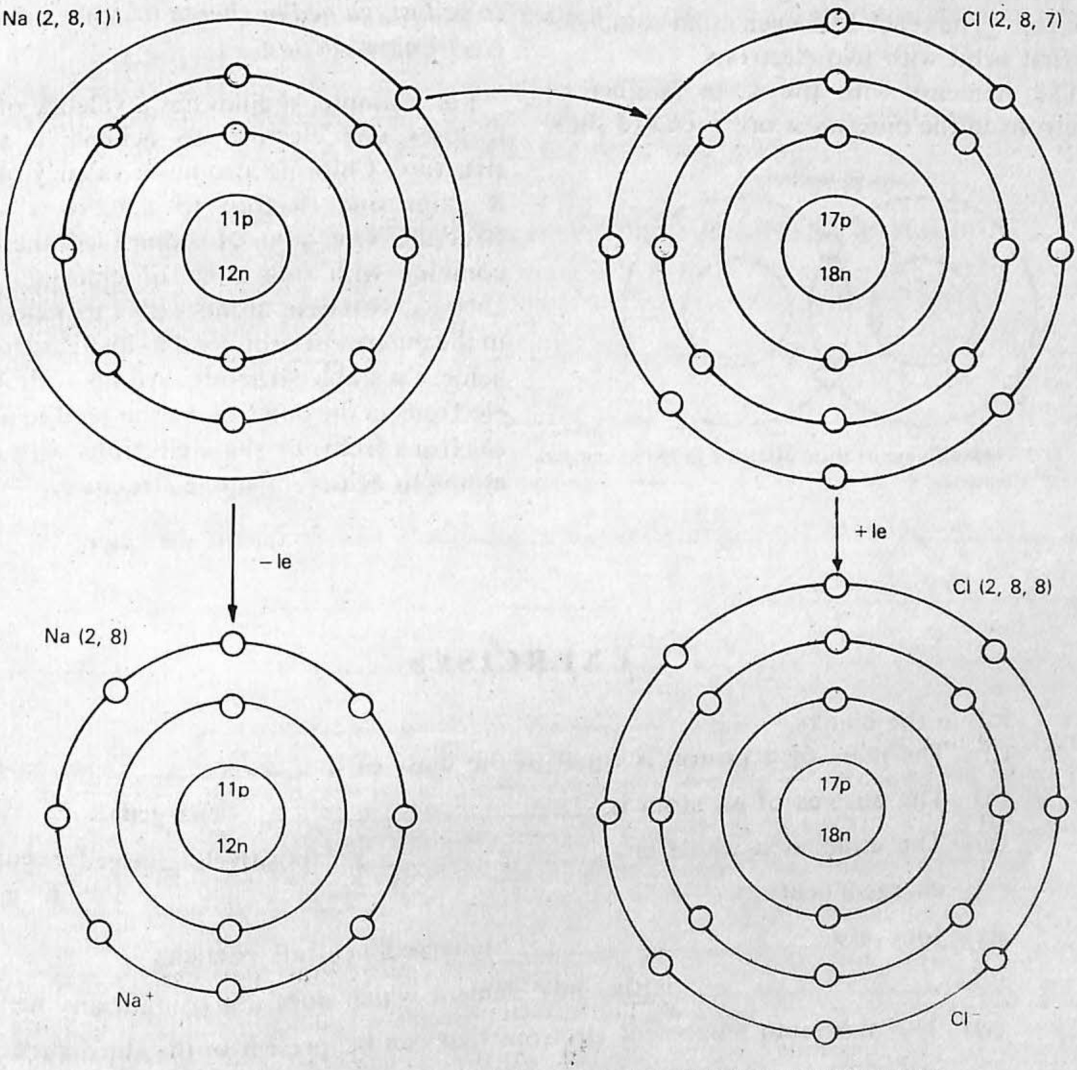


Fig. 4.10 Sodium reacts with chlorine by losing one electron which is accepted by chlorine.

Sodium (2, 8) (stable electronic configuration of neon)
Chlorine (2, 8, 8) (stable electronic configuration of argon).
Both of these are stable with 8 electrons in their outermost orbits.

When sodium loses one electron it becomes positively charged. When chlorine gains one electron it becomes negatively charged. These charged particles are called **ions.**

Atoms can also combine by sharing electrons. Two hydrogen atoms combine to form a molecule of hydrogen by sharing their electrons. Thus each hydrogen atom completes its first orbit with two electrons.

The elements with the same number of electrons in the outermost orbit tend to show similar chemical properties. For example lithium (2, 1), sodium (2, 8, 1) and potassium (2, 8, 8, 1) have similar chemical properties. This shows the importance of electrons in determining the chemical nature of an element.

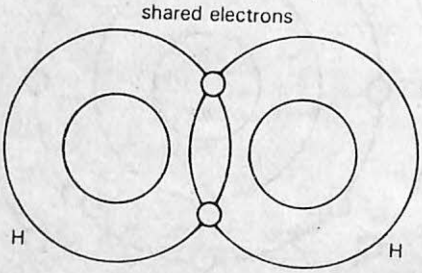


Fig. 4.11 Hydrogen atoms share electrons to form hydrogen molecule.

4.8 Valency

We have seen in Class 7 that when atoms combine to form molecules, the number of atoms which join together depend on their combining power or valency. *Valency represents the number of electrons which need to be lost, gained or shared to achieve a stable electronic structure.*

For example, sodium has a valency of 1 as it loses one electron to achieve a stable structure. Chlorine also has a valency of 1 as it gains one electron to achieve a stable structure. One atom of sodium will therefore combine with one atom of chlorine. In a chemical reaction, atoms with 1 to 3 electrons in the outermost orbit tend to lose electrons to achieve a stable structure. Atoms with 4 to 7 electrons in the outermost orbit tend to accept electrons from, or share electrons with other atoms to achieve a stable structure.

EXERCISES

1. Fill in the blanks.
   (i) The mass of a proton is equal to the mass of a ________________ atom.
   (ii) The nucleus of an atom is________________________ charged.
   (iii) The atom as a whole is ________________ (positively charged/negatively charged/neutral).
   (iv) Ions are________________________(charged/neutral) particles.
   (v) ________________ is the only element which does not contain any neutron.
   (vi) The maximum number of electrons that can be present in the third shell in an atom is ________________.

(vii) The sub-atomic particles__________ and __________ are equal in mass.

(viii) Atoms of the same element having same number of protons, but different number of neutrons are called__________of the element.

2. Indicate true or false.
   - (i) Hydrogen and its isotopes do not contain any neutrons in their nucleus.
   - (ii) An atom is a solid mass.
   - (iii) Two atoms having the same atomic number can exhibit different chemical properties if they have different mass numbers.
   - (iv) An atom having the electronic configuration 1, 8, 3 can exist.
   - (v) Helium is a non reactive element.
   - (vi) Since mass number = number of protons + number of neutrons, mass number is always a whole number.
   - (vii) Relative atomic mass and mass number are the same.
   - (viii) Ions are charged particles formed when atoms lose or gain electrons.
   - (ix) Valency = number of electrons in the outermost orbit.

3. In what important ways does the Bohr model differ from the solar system? __________

4. (a) What are isotopes? List the isotopes of (i) hydrogen (ii) chlorine. __________

   (b) Explain why the relative atomic mass of chlorine is 35.5. __________

5. (a) What is the tendency of any atom during a chemical reaction?

(b) Explain the above with reference to the reaction between sodium and chlorine.

6. Complete the following table.

Element	*Atomic number*	*Electronic configuration*
Oxygen	8	
Carbon	6	
Sulphur	16	
Potassium	19	
Manganese	25	

7. Complete the following table.

Element	*No. of neutrons*	*No. of protons*	*No. of electrons*	*Atomic number*	*Mass number*
Beryllium	5	4			
Fluorine	10		9		
Sodium				11	23
Magnesium	12			12	
Argon		18			40

5

Nuclear Chemistry

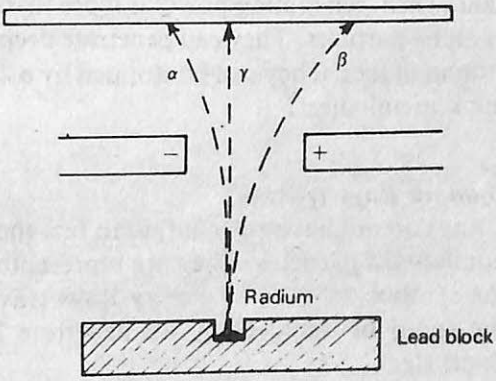
5.1 Radioactivity

In 1896, a French scientist, Henry Becquerel. accidently discovered a strange natural phenomenon. He happened to place a small quantity of a mineral called 'pitch blende' (a compound of the element uranium) in his desk drawer. A photographic plate shielded with a cover happened to be in his desk. A key was accidently left between the photographic plate and the pitch blende. Later, when he developed the photographic plate, he noticed that an impression of the key was left on the photograph. He concluded that some mysterious radiations were emitted by the pitch blende, capable of penetrating the cover of the photographic plate.

In subsequent years, many scientists worked to understand these mysterious radiations. They named the substances which release such radiations as **radioactive** substances. The phenomenon was called **radioactivity.** Later Marie Curie and Pierre Curie discovered radium, which is far more radioactive than uranium.

It was discovered that these radiations were capable of penetrating most known materials including metals but were unable to penetrate lead. Continued exposure to radioactivity is harmful to the human body. Unfortunately, this was realised only after some scientists working on radioactive materials died or suffered badly.

It was discovered that there were three types of radiations coming from radioactive substances. One type of radiation contained positively charged particles, another contained negatively charged particles and the third was neutral. These three types of radiations were



α particles attracted towards negative plate
β particles attracted towards positive plate
γ particles not attracted

Fig. 5.1 Effect of electric field on α, β and γ radiations.

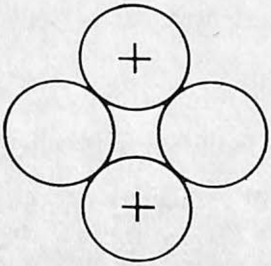


Fig. 5.2 An α particle

named **alpha, beta** and **gamma** rays, respectively.

Subsequent experiments revealed many distinct characteristics of these radiations.

Alpha Rays (α-Rays)

α-Rays consist of particles with positive charge, travelling at a speed of 3200 to 32,000 km per second. Each particle consists of 2 protons and 2 neutrons, similar to the nucleus of helium gas. Their penetrating power is low. They can penetrate human skin or paper.

Beta Rays (β-Rays)

β-Rays are a stream of electrons with negative charge. They are produced by the splitting of a proton into a neutron and an electron. β-rays travel at a speed of 96,000 to 250,000 km per second which is slightly less than the speed of light. Their penetrating power is more than that of alpha particles. They can penetrate deep into human tissues. They can be stopped by a 2 mm thick metal sheet.

Gamma Rays (γ-Rays)

γ-Rays do not have any charge. In fact they do not contain particles. They are represented by the symbol $\rightsquigarrow$. γ-Rays travel at the speed of light. They can penetrate 2.5m thick steel.

All three radioactive rays emanate from within the nucleus of atoms. Due to radioactivity, the nucleus of an element undergoes a change and hence the element changes to a new one. It should be remembered that radioactivity is not an ordinary chemical change which involves electrons in the orbits of the atom. It is a nuclear change. A nuclear change occurs due to changes in the nucleus of an atom. It causes the atom to transform itself to the atom of a new element.

When a nucleus emits an α-particle it loses 2 protons (apart from two neutrons) and consequently the atomic number of the remaining nucleus reduces by 2. The mass number reduces by 4. A uranium ($^{238}_{92}U$) nucleus, after emitting an α-particle is transformed to thorium ($^{234}_{90}Th$) nucleus. Similarly radium will change to radon.

A β-particle is emitted after a neutron decomposes to a proton and an electron, and this electron leaves the nucleus. The atomic number of the nucleus increases by one unit and the atom acquires the identity of a new element. By emission of a β-particle, uranium transforms to neptunium, which in turn transforms to plutonium.

Gamma radiation does not, by itself, result in a nuclear change. γ-Radiation usually accompanies emission of α or β particles.

Half Life

Radioactive atoms are unstable. They are constantly emitting radiations and changing to new atoms. This process is called **radioactive decay.** The rate of decay is expressed in terms of the half life of the element. Half life is defined as the *time required for half of a given quantity of a radioactive element to change to another element.* The half life of different radioactive elements may vary from few billionths of a second to millions of years. Half life of radium is 1590 years. This means that 1g of radium taken today would reduce to 0.5g in 1590 years from now, and to 0.25 (half of 0.5) in another 1590 years, and so on.

It is observed that all nuclei with atomic number equal to 83 or more are unstable and hence radioactive. All these elements tend to decay till they become lead ($_{82}Pb$) which is highly stable. Besides, many lighter elements such as carbon, also have radioactive isotopes.

5.2 Fission

Nuclear changes can be brought about artificially also. *The splitting of a heavy nucleus to form two smaller nuclei of about equal size is known as fission.* Fission can be brought about by bombarding the nucleus of a radioactive element such as uranium with high velocity neutrons. The nucleus of uranium breaks up to form atoms of two new elements—barium and krypton. During this process a large amount of energy is released and more high velocity neutrons are produced.

These neutrons cause the fission of other nuclei. Thus a chain reaction is set up and a large quantity of energy is released from a small quantity of radioactive material. This principle of fission is used in the **atomic bomb**.

However, if this chain reaction is controlled the explosion can be prevented and the energy released can be used. This is the principle of an **atomic reactor.**

The energy released can, for example, be used to generate electricity. Several such nuclear power stations are in operation in our country.

5.3 Fusion

In another type of nuclear reaction, two or more atomic nuclei can be fused together to form one larger nucleus. In this process also a very large amount of energy is liberated. Such nuclear reactions are known as **fusion.** To start a fusion reaction a large amount of energy is required. In one such fusion reaction hydrogen nuclei combine to form a helium nucleus. This reaction, in fact, is the source of the energy of the sun and stars, where a very large number of such reactions occur every second. The hydrogen bomb, which is much more powerful than an atomic bomb, also employs this reaction.

5.4 Uses of Radioactivity

1. Determination of Age

Knowing the ratio of $^{238}_{92}U$ to $^{206}_{82}Pb$ present in rocks containing uranium mineral and knowing the half life of uranium, the age of the earth has been estimated as 4 to 5 billion years.

Carbon dating is used for determining the age of dead plants and other organic materials. This is based on the principle that the carbon dioxide absorbed by the plant while alive, contained ^{12}C and ^{14}C in a known definite

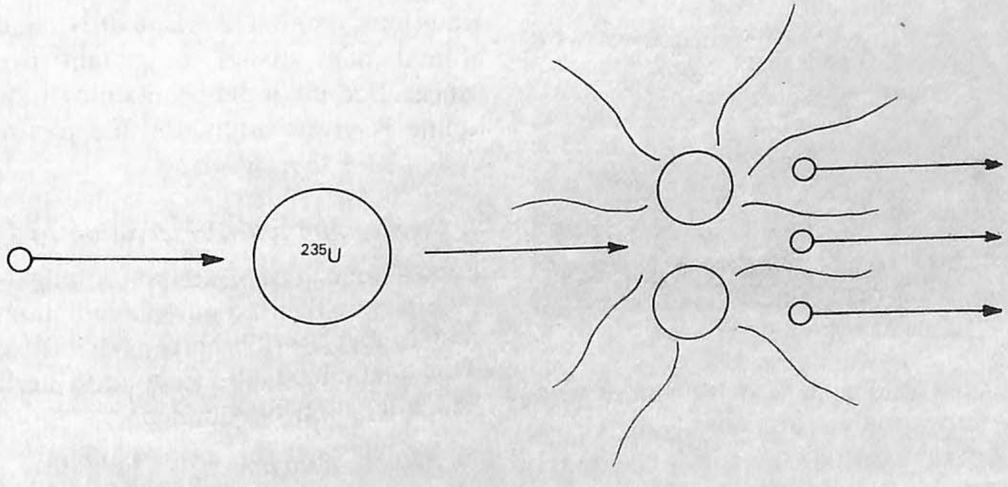


Fig. 5.3. Fission of urainium – 235 atom

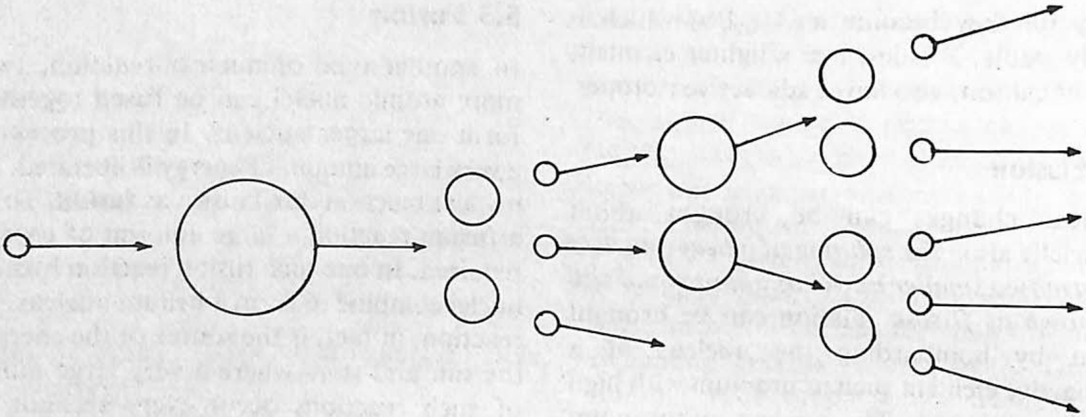
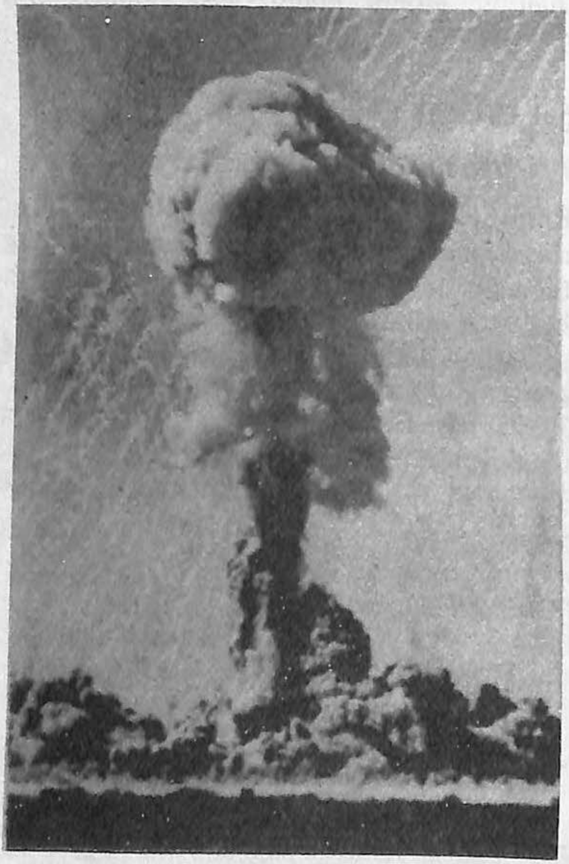


Fig. 5.4(a) Chain reaction in an atomic bomb.



(b) An atomic bomb explosion.

ratio. In the dead plant there is no more intake of carbon dioxide and the existing ^{14}C continues to decay. By measuring the ratio of ^{12}C and ^{14}C in a dead plant or other organic specimen, the age of the specimen can be determined. Carbon dating has been used, for example, to determine the age of Egyptian mummies.

2. As Tracers

By introducing a radioactive element into the body of humans, animals or plants the tracing of the path of the element can be studied to understand the complex internal processes. A good example is introduction of carbon dioxide with radioactive ^{14}C into a plant to trace the course of photosynthesis.

3. Medical Applications

Raditions emitted by Cobalt-60 is directed at malignant tissues to contain growth of cancer. Sodium iodide containing radioactive iodine is given orally for the treatment of underactive thyroid gland.

4. Sterilisation and Preservation of Food

Food items can be preserved for long durations by exposing them to nuclear radiations which destroy microorganisms responsible for decay. This method can also be used to sterilise, for example, surgical equipment.

5. Atomic Reactors

Controlled nuclear fission is done in atomic reactors. The intense heat released is used to generate steam at high pressure. This steam is used to turn turbines to generate electricity.

Notwithstanding the good uses of radioactivity, too much exposure to these radiations is harmful to men and animals. Excessive exposure can destroy the blood cells, and cause damage to the central nervous system.

EXERCISES

1. Fill in the blanks:

(i) ______________________ discovered radioactivity.

(ii) Particles of α-rays are similar to the nucleus of ______________________

(iii) ______________________ rays are nothing but streams of electrons.

(iv) If an α-particle is emitted from an element its atomic number is reduced by ______________________.

(v) The rate of decay of a radioactive element is expressed in terms of ______________.

(vi) Atomic reactors work on the principle of ______________________.

(vii) ______________ isotope of carbon is used to determine the age of dead plants.

(viii) ______________________ is used for the treatment of cancer.

(ix) Goitre, a disease caused by the malfunctioning of thyroid gland, is treated using ______________________.

2. Explain briefly:

(i) Nuclear fission. ____

____

(ii) Half life period. ____

____

(iii) Nuclear fusion. ____

____

(iv) Radioactive elements as tracers. ____

____

(v) Ill effects of excessive exposure to radioactivity. ____

____

3. (i) Compare the properties of α, β, γ rays.

(ii) How does radioactivity help in the preservation of food articles?

(iii) What is the function of atomic reactors in an atomic power station?

(iv) An element $^{242}_{94}X$ emits one alpha particle and one beta particle. What will be the atomic number and mass number of the new element formed?

4. State true or false.

(i) γ-rays are the most penetrating of all radiations emitted during radioactive decay.
(ii) γ-rays consist of neutral particles.
(iii) γ-rays travel with the speed of sound.
(iv) Radioactive decay results in a chemical reaction.
(v) Half life of a radioactive element is 1 day. 5 g of the element are taken. It will decay to 2.5 g in one day and to 0 g in 2 days.
(vi) A hydrogen bomb employs fission reaction to liberate large amounts of energy.
(vii) A nuclear reactor employs controlled fusion reaction to generate electricity.

6

Electricity and Chemical Changes

You have studied in physics that electricity passes readily through metallic elements. Copper and aluminium cables are used to bring electricity to our houses. They get heated by the passage of electricity through them, but they do not undergo any chemical change. Such substances are called **conductors.**

Non-metallic elements are bad conductors of electricity. Graphite, a form of carbon is, however, a good conductor of electricity.

6.1 Conduction of Electricity through Compounds

Activity 1: Connect two platinum wires to a 6 volt battery, with an ammeter reading upto 2 amperes in series (Fig. 6.1). Take distilled water in a beaker. Dip the platinum wires into it and note the current flowing in the ammeter. You will notice that the ammeter does not record any current. This shows that distilled water does not conduct electricity. Repeat the experiment taking the following liquids separately in beakers:

(a) alcohol
(b) dilute hydrochloric acid
(c) dilute sulphuric acid
(d) sugar solution in water
(e) sodium chloride solution in water
(f) copper sulphate solution in water
(g) sodium hydroxide solution in water

Remember to wash the platinum wires with distilled water after they have been dipped into one solution and before dipping into another solution.

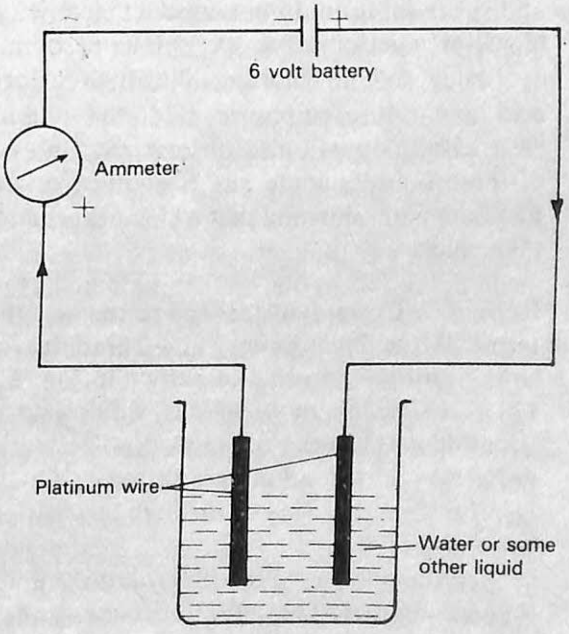


Fig. 6.1 Conduction of electricity by liquids.

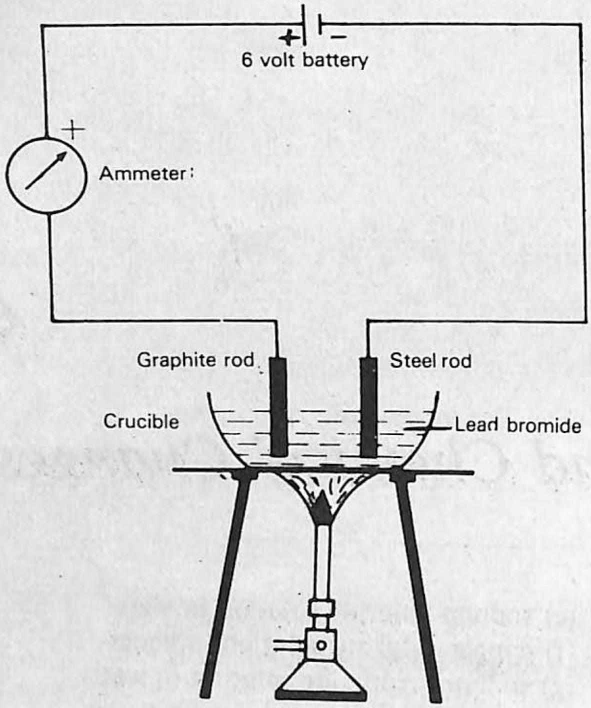


Fig. 6.2 Conduction of electricity through molten lead bromide.

You will observe that distilled water, alcohol and sugar solution do not conduct electricity. However, solutions of copper sulphate, sodium hydroxide, sodium chloride, dilute hydrochloric acid and dilute sulphuric acid do conduct electricity. You will also observe that in case of these liquids, some gas is evolved at the platinum wire, showing that a chemical change takes place.

Activity 2: Connect a steel rod to the negative terminal of a 6 volt battery and a graphite rod to the positive terminal as shown in Fig. 6.2. Fill a crucible, about two-thirds, with solid lead bromide and dip the rods into it. There is no deflection in the ammeter, showing that no electric current flows through solid lead bromide.

Melt the solid lead bromide by heating it with a bunsen burner. The deflection in the ammeter shows the flow of electric current.

Lead bromide is soluble in water. If rods connected to the battery are dipped into a solution of lead bromide, the flow of electric current is shown by the deflection of the needle in the ammeter.

You will observe that when electricity flows through molten lead bromide, red vapours of bromine are liberated around the carbon rod connected to the positive terminal, and metal lead is deposited around the steel rod connected to the negative terminal of the battery. This clearly indicates the chemical decomposition of the compound which conducts electricity.

6.2 Electrolyte and Electrolysis

As far as the conduction of electricity through compounds is concerned, they can be put into two definite classes. Compounds such as acids, bases and salts conduct electricity and are called **electrolytes.** Compounds such as alcohol and sugar which do not conduct electricity are called **non-electrolytes.** Water conducts electricity at a very high voltage.

Thus electrolytes are compounds which conduct electricity, when in molten state or when dissolved in water, and undergo chemical decomposition by the flow of electric current. The process of conduction of electricity through an electrolyte, which undergoes a chemical decomposition by the flow of electricity, is called **electrolysis.**

6.3 Electrodes

The metal or carbon (graphite) rods used in electrolysis are called electrodes. The electrode connected to the **positive terminal** of the battery is called the **anode** while the electrode connected to the **negative terminal** is called the **cathode.**

Platinum and carbon electrodes are generally used in electrolysis because they are non-reactive and hence they react neither with the electrolyte nor with the products of electrolysis.

6.4 Atoms and Ions

Electrolytes contain charged particles, with positive and negative charges, which are attracted by electrodes having opposite charges. Positively charged particles are attracted by the cathode and the negatively charged particles by the anode. These charged particles are responsible for the conduction of electricity.

Charged atoms or groups of atoms are called **ions.** In the electrolysis of molten lead bromide, metallic lead is deposited on the cathode (negative plate) and the non-metal bromine is collected around the anode (positive plate). In the compound lead bromide, the lead must, therefore, be a positive ion and bromide must be a negative ion. Experiments on electrolysis of other compounds show that all metallic and hydrogen ions are positively charged and non-metallic ions are negatively charged.

The positively charged metallic and hydrogen ions are called **cations.** *These are discharged at the cathode (negative electrode) during electrolysis. The non-metallic ions are called* **anions.** *During electrolysis these are discharged at the anode (positive electrode).*

The ions of an element are completely different from their atoms. For example, an atom of sodium is electrically neutral whereas an ion of the element carries a unit positive charge. When sodium metal is added to water, there is a vigorous reaction producing hydrogen gas and sodium hydroxide solution. The sodium ions present in sodium chloride do not react at all with water. Similarly an atom of copper is electrically neutral whereas an ion of copper carries two units of positive charge.

6.5 Charge on Ions

In electrolytes, the ions are responsible for the conduction of electricity. If an electrolyte contains a large number of ions, it is called a **strong electrolyte** and is a better conductor of electricity. On the other hand a **weak electrolyte** contains few ions and is a poor conductor of electricity. Water, ammonium hydroxide and acetic acid are weak electrolytes. Hydrochloric acid, sulphuric acid, nitric acid, the salts of these acids and caustic alkalis are strong electrolytes.

The charge on the ions can be found by applying Faraday's laws of electrolysis, which you will study in higher classes. *The charge on an ion is the same as the valency of the element or radical.* When an atom loses an electron from the outermost orbit it becomes a positively charged ion. If it gains an electron it transforms to a negatively charged ion. The charge on an ion is indicated on the right of the symbol at the top. Thus CO_3^{2-} indicates that the carbonate ion carries two negative charges. The sodium ion is Na^+, copper ion is Cu^{2+}, aluminium ion is Al^{3+} and ammonium ion is NH_4^+. Some of the non-metallic ions are: chloride ion Cl^-, sulphate ion SO_4^{2-} and phosphate ion PO_4^{3-}.

6.6 Electrolysis of Acidified Water

Activity 3: Different types of apparatus have been used to study the effect of electrolysis of water which is acidified with dilute sulphuric acid. Either platinum or carbon electrodes can be used. In the apparatus shown in Fig. 6.3, two carbon electrodes pass through a vessel. Connect the electrodes with a 6-volt battery in series with an ammeter. Fill the vessel with water containing a few cm³ of dilute sulphuric acid. Allow the current to pass for at least 15 minutes. The gases evolved are collected in the test tubes by displacing water.

Compare the volumes of the two gases. Test the gas liberated at the anode with a glowing splint and the gas liberated at the cathode with a burning splint. You will notice that the gas evolved at the anode is oxygen and the gas liberated at the cathode is hydrogen. Oxygen and hydrogen are in the ratio of 1:2 by volume. The products indicate that water has been

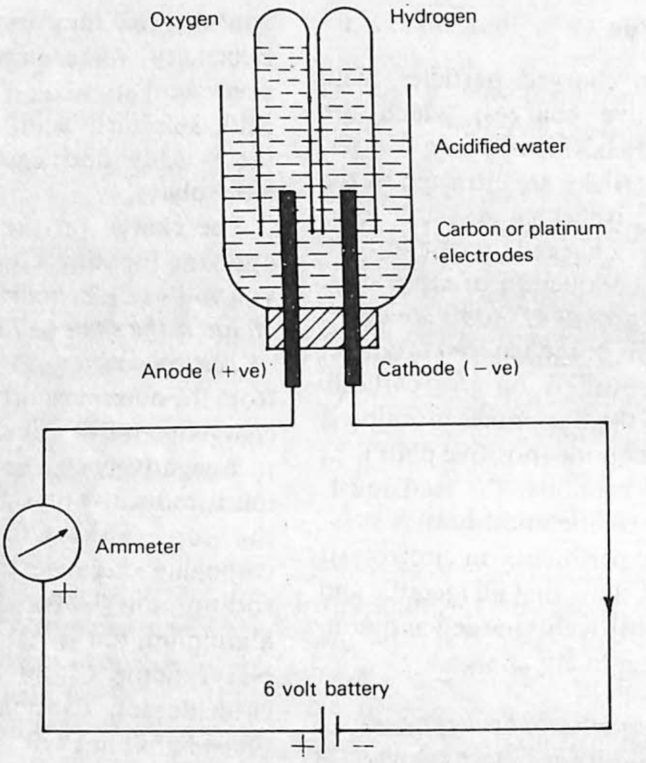


Fig. 6.3 Electrolysis of acidified water.

decomposed into oxygen and hydrogen during electrolysis.

$2H_2O \rightarrow 2H_2 + O_2$

6.7 Electrolysis of Copper Chloride Solution

Activity 4: Fill a U-tube with 30 per cent solution of copper chloride in water. Dip two platinum electrodes into the solution and connect the electrodes to a 10-volt battery and an ammeter (Fig. 6.4).

You will observe that a reddish-brown coating of copper metal is deposited at the cathode. There is an evolution of a greenish yellow gas around the anode, which comes out of the U-tube and can bleach moist litmus paper. This gas is chlorine.

$CuCl_2 \rightarrow Cu + Cl_2$

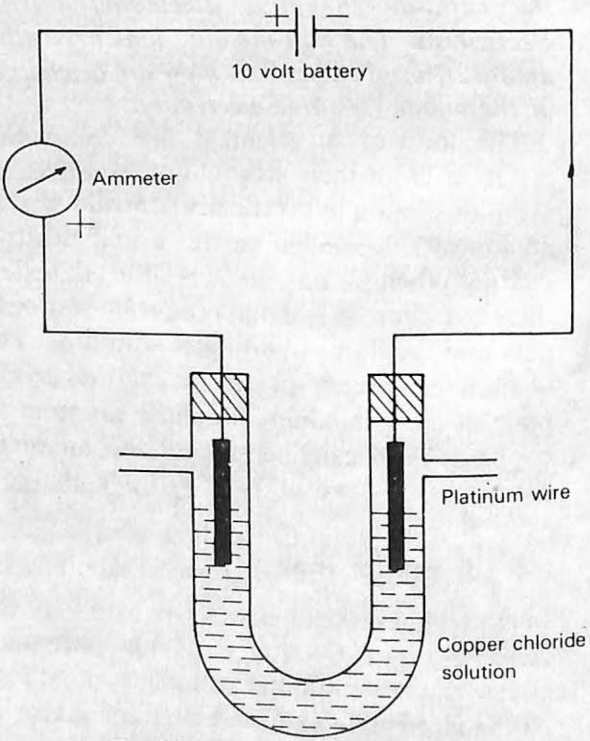


Fig. 6.4 Electrolysis of copper chloride.

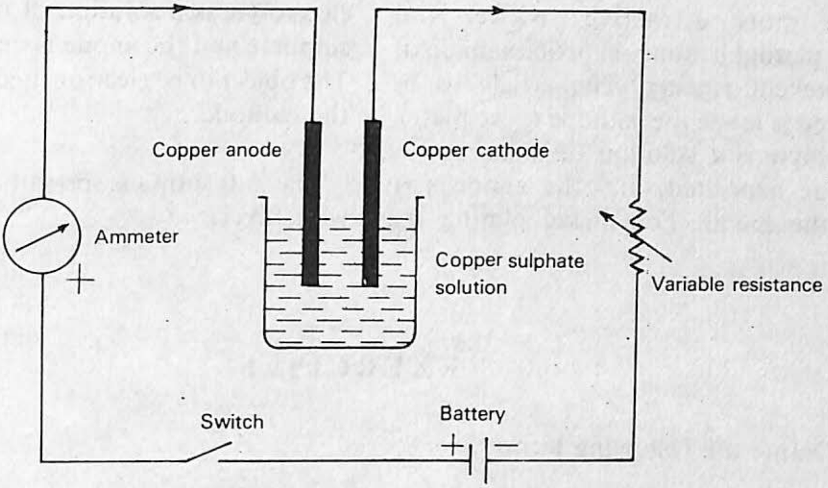


Fig. 6.5 Electrolysis of copper sulphate using copper electrodes.

6.8 Electrolysis of Copper Sulphate Solution

When electricity is passed through copper sulphate solution using both electrodes of copper (Fig. 6.5), the electrolysis does not result in the decomposition of copper sulphate. Copper ions from the copper sulphate solution are deposited on the cathode. At the anode, copper atoms change into copper ions and go into the solution. The reactions at the two electrodes may be written as below:

At cathode: Cu^{2+} changes to Cu

At anode: Cu changes to Cu^{2+}

In effect, copper is transferred from the anode to the cathode. Thus the cathode becomes thicker and the anode becomes thinner as the electrolysis proceeds. This principle is applied in the refining of copper when thin sheets of pure copper are made the cathode and thick sheets of impure copper are made the anode. Pure copper gets transferred from the anode to the cathode and the impurities settle at the bottom of the tank.

6.9 Applications of Electrolysis

1. *Extraction of metals:* More reactive metals such as sodium could be extracted from minerals only when electrolysis became known to the scientists.

2. *Refining of metals:* A large number of metals are obtained in pure form by electrolysis. Copper refining has been explained earlier.

3. *Electroplating:* It is the process of depositing a thin and compact film of a non-reactive metal on an article made of another metal. Silver and gold plating is done to make

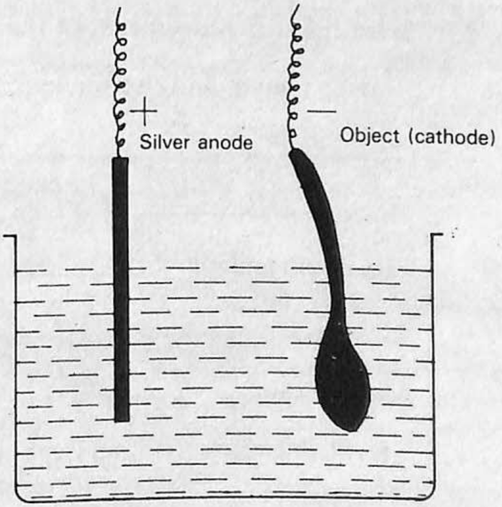


Fig. 6.6 Silver plating by electrolysis.

an article more attractive. Nickel and chromium plating is done on articles made of steel to prevent rusting. The article to be electroplated is made the cathode (– ve plate). The electrolyte is a solution of a salt of the metal to be deposited, and the anode(+ ve plate) is the metal. For nickel plating the electrolyte is a solution of nickel ammonium sulphate and the anode is made of nickel rod. The object to be electroplated by nickel is made the cathode.

Fig. 6.6 shows a spoon being electroplated with silver.

EXERCISES

1. Define the following terms.

(a) Electrolyte ______________________________

____

(b) Ion ______________________________

____

(c) Cation ______________________________

____

(d) Anion ______________________________

____

2. State the difference between the following.

(a) Conductor and electrolyte ______________________________

____

____

(b) Atom and ion ______________________________

____

____

(c) Strong electrolyte and weak electrolyte ______________________________

____

____

3. Fill in the missing words.

Molten lead bromide consists of ______________ ions and ______________ ions. The ion of______________ is called ______________ and the ion of______________ is called ______________________ When electricity is passed through molten lead bromide the ________________ions are deposited on the ________________ and the________________ ions are deposited on the________________ The electrode connected to the positive terminal is called____________________ and the electrode connected to the negative terminal is called____________________ .

4. (a) State two industrial uses of electrolysis.

____

____

____

(b) Draw a labelled diagram for electroplating an object with copper.

5 (a) Explain what you would observe when electricity is passed through water containing a little sulphuric acid. ______________________

____

____

____

____

____

(b) How is pure copper obtained from impure copper? ______________

____

____

____

____

6. (a) Name an acid which is a weak electrolyte. ______________________

____

(b) Give one example each of an ion containing more than one element, and an ion having a positive charge. ______________________________

____

7. Explain the following:

(a) Solid sodium chloride does not conduct electricity whereas molten sodium chloride does. ______________________________

____

(b) An aqueous solution of acetic acid is a poor conductor of electricity whereas dilute sulphuric acid is a good conductor. ______________________________

____

____

8. (a) Write down the ions present, along with the charges on them, in sodium carbonate, ammonium sulphate and sulphuric acid. ______________________________

____

____

(b) State what would happen when electricity is passed through molten magnesium chloride. ______________________________

____

____

____

____

7

Metals and Alloys

7.1 Minerals, Ores and Rocks

The earth is a treasure house of chemical materials. These materials contain elements and compounds of elements called **minerals.** These minerals are mixed with impurities such as clay and sand. The mixture of impurities and minerals is called **ore.** Metals are extracted from ores. The unwanted impurities, e.g. rocky matter and sand, called **gangue,** are removed. This process is called **concentration** of ore. The minerals, free from impurities, are used for the extraction of metals. **Haematite,** an ore containing the mineral iron (III) oxide is used to obtain iron and steel. **Bauxite** contains the mineral aluminium oxide and is used for extracting aluminium. **Copper pyrites,** containing sulphides of copper and iron, is used to obtain copper. Some unreactive metals, such as gold, are found in an uncombined state in nature. Sulphur also occurs in nature in the free state.

Rocks contain one or more minerals with or without clay and sand. Calcium carbonate is abundantly found in the rock formations such as chalk, limestone and marble. These are formed when shells and other hard parts of dead animals are subjected to high temperature and pressure inside the earth's crust.

7.2 Metals

Metals are of great importance to man. Metals such as iron, aluminium, copper and lead find various applications in our every day life. There are many other metals we do not see around us, which are also important, e.g. uranium, a very rare metal, is the basic material for generation of atomic energy.

The metallic resources of a nation are of immense value. Silver and gold were the first metals discovered. The discovery of bronze, an alloy of copper and tin, resulted in the *Bronze Age* in history. This period was followed by the *Iron Age*, with the discovery of superior quality of iron. A number of uses of metals are based on their physical properties such as malleability, ductility, conduction of electricity and high tensile strength.

7.3 Occurrence of Metals

Metals which are not chemically reactive, e.g. gold and silver, are found in the uncombined state called the **native state.** This explains why gold and silver were discovered early.

Minerals of metals occur generally as oxides, sulphides, carbonates and chlorides. These minerals are mixed with gravel, sand and other earthy impurities from which they must be separated by physical or chemical processes.

The minerals present in the earth's crust are decreasing as they are being used in extraction of metals. The National Institute of Oceanography has completed a survey discovering good polymetallic nodule fields in the Indian Ocean. It is expected that in years to come extraction of some useful metals from the nodules should be possible on a commercial scale.

7.4 Extraction of Metals

Some of the principles for the extraction of metals from their ores are as follows:

1. Reactive metals such as potassium, sodium, calcium and magnesium are extracted by the *electrolysis of their molten chlorides.*

For example, sodium chloride whose melting point is 800°C is mixed with calcium chloride which brings down its melting point to about 600°C. When electricity is passed through the molten mixture of the two salts, sodium, being a positively charged ion, is collected around a steel cathode. Chloride ions are discharged on a graphite anode and form molecules of chlorine which are collected as greenish-yellow chlorine gas. Thus:

At the cathode: Sodium ions change to sodium atoms.

At the anode: Chloride ions change to chlorine gas.

2. Less reactive metals are extracted by a method which generally involves the following steps:

(a) *Concentration of the ore:* This step leads to the removal of earthy impurities.

(b) *Roasting of the ore:* The purpose of roasting, or heating in air, is to change carbonate and sulphide ores of metals to oxides.

(c) *Reduction:* The oxide is reduced to the metal by removal of oxygen. The common reducing agent used is coke, a form of carbon. The reduction is carried out in a furnace whose temperature is about 1000-1500°C. For example, the ore of iron called haematite is mixed with coke and heated in a blast furnace. The coke first burns in the furnace to produce carbon dioxide which changes to carbon monoxide.

$$\underset{\text{coke}}{C} + O_2 \rightarrow \underset{\text{carbon dioxide}}{CO_2}$$

$$C + CO_2 \rightarrow \underset{\text{carbon monoxide}}{2CO}$$

The carbon monoxide reduces the haematite to form iron.

$$3CO + \underset{\text{haematite}}{Fe_2O_3} \rightarrow \underset{\text{iron}}{2Fe} + 3CO_2$$

The ore of zinc which contains zinc carbonate or sulphide when roasted in air changes to zinc oxide. This is mixed with coke and heated to give the metal.

$$\underset{\text{zinc oxide}}{ZnO} + \underset{\text{coke}}{C} \rightarrow \underset{\text{zinc}}{Zn} + \underset{\text{carbon monoxide}}{CO}$$

(d) *Refining of the metal:* This is often done by electrolysis but other processes are also in use.

7.5 Physical Properties of Metals

1. Metals are solids at room temperature. Mercury is, however, a liquid metal and gallium metal melts at body temperature.

2. Metals possess a lustre called **metallic lustre**, when freshly cut. But they may tarnish when exposed to air.

3. Metals generally have a high density. Sodium and potassium are examples of some light metals which are even lighter than water.

4. Metals are generally hard. Sodium and potassium are, however, very soft and can be cut by a knife.

5. Metals can be generally beaten into sheets. This property is called **malleability**. Sheets of iron are used for making car bodies and machines. Thin sheets of aluminium, called aluminium foil, are used to wrap chocolates and medicines.

6. Some metals can be drawn into thin wires. This property is called **ductility**. Wires of aluminium and copper are used for conduction of electricity.

7. Metals are good conductors of heat and electricity. Copper wire is widely used in house wiring and in electrical machines and heating appliances such as geysers.

7.6 Chemical Properties of Metals

1. A large number of metals react with water at different temperatures.

(a) Sodium and potassium react vigorously with cold water forming their hydroxides, which are soluble, and hydrogen.

$$\underset{\text{sodium}}{2Na} + 2H_2O \rightarrow \underset{\text{sodium hydroxide}}{2NaOH} + H_2$$

(b) Calcium reacts with cold water slowly forming calcium hydroxide which turns water turbid and hydrogen.

$$\underset{\text{calcium}}{Ca} + 2H_2O \rightarrow \underset{\text{calcium hydroxide}}{Ca(OH)_2} + H_2$$

(c) Magnesium will react with boiling water to give magnesium oxide and hydrogen.

$$\underset{\text{magnesium}}{Mg} + H_2O \rightarrow \underset{\text{magnesium oxide}}{MgO} + H_2$$

(d) Aluminium, even though a reactive metal, does not react with water because of a compact and inert layer of aluminium oxide on the metallic surface.

Aluminium, which has been freed from its oxide layer readily reacts with hot water producing aluminium hydroxide and hydrogen.

$$\underset{\text{aluminium}}{2Al} + 6H_2O \rightarrow \underset{\text{aluminium hydroxide}}{2Al(OH)_3} + 3H_2$$

It is because of the oxide layer of aluminium that the cooking utensils of aluminium are not attacked by water.

(e) Red hot zinc reacts with boiling water to give zinc oxide and hydrogen.

$$\underset{\text{zinc}}{Zn} + H_2O \rightarrow \underset{\text{zinc oxide}}{ZnO} + H_2$$

(f) Red hot iron reacts with steam to give tri-iron tetroxide and hydrogen.

$$\underset{\text{iron}}{3Fe} + 4H_2O \rightarrow \underset{\textit{tri-iron tetroxide}}{Fe_3O_4} + 4H_2$$

At very high temperature lead will react slightly to give its oxide. Copper does not react with water at any temperature.

The reaction of metals with water shows that different metals have different reactivities. Reaction of dilute hydrochloric acid with metals also shows different degrees of reactivities of metals. Metals can, therefore, be arranged in the decreasing order of their reactivity. This arrangement is called **reactivity series of metals** and is as follows:

K, Na, Ca, Mg, Al, Zn, Fe, Pb, (H), Cu, Hg, Ag, Au, Pt.

Hydrogen, although a non-metal, has been included in the series because in some ways it shows metallic properties, e.g. it forms positive ions like metals.

2. Metals combine with oxygen to form oxides, which are generally *basic oxides*.

$$\underset{\text{magnesium}}{2\,Mg} + O_2 \rightarrow \underset{\text{magnesium oxide}}{2MgO}$$

$$\underset{\text{sodium}}{4Na} + O_2 \rightarrow \underset{\text{sodium oxide}}{2Na_2O}$$

The metallic oxides neutralise an acid forming a salt and water.

$$\underset{\text{sodium oxide}}{Na_2O} + \underset{\text{hydrochloric acid}}{2HCl} \rightarrow \underset{\text{sodium chloride}}{2NaCl} + H_2O$$

3. Dilute hydrochloric acid and dilute sulphuric acid react with metals above hydrogen in the

reactivity series to give a salt and hydrogen.

Fe	+	H_2SO_4	→	$FeSO_4$	+	H_2
iron		dil. sulphuric acid		iron (II) sulphate		

4. A more reactive metal displaces a less reactive metal from its compound.

Zn	+	$CuSO_4$	→	$ZnSO_4$	+	Cu
zinc		copper (II) sulphate		zinc sulphate		copper

5. Oxides of metals do not decompose on heating except oxides of mercury and silver which produce the metal and oxygen.

$2HgO$	→	$2Hg$	+	O_2
mercury(II) oxide		mercury		

7.7 Classification of Metals

There are many ways of classifying metals. Some of these are as follows.

1. *Classification based on reactivity:* You have seen that metals differ in their reactivity. This can be shown by the reaction of metals with water and oxygen of the air. On this basis we can arrange the metals in a series called **reactivity series.** This series has been given earlier.

2. *Classification based on other properties:* Metals can be divided into two groups:

(a) Iron and steel which form the backbone of industries, and are called ferrous metals.

(b) Non-ferrous metals which can be further subdivided into the following groups:

  (i) Light metals, e.g. potassium, sodium, magnesium and aluminium.

  (ii) Heavy metals, e.g. lead, mercury, copper, silver, gold and platinum.

  (iii) Protective metals, e.g. chromium, nickel, zinc and tin.

7.8 Alloys

Alloys are mixtures of metals prepared by melting and mixing the component metals in definite proportions by weight.

Alloys are prepared to obtain properties which are not possessed by component metals. The properties which change in alloying metals are hardness, malleability, resistance to *corrosion* (which means chemical attack by atmosphere, water and soil) and melting point. *Amalgam* is an alloy of a metal with mercury.

The composition, properties and uses of some of the common alloys are given in Table 7.1.

7.9 Corrosion of Metals

Most metals are attacked chemically by the atmosphere, by soil and by water. This chemical attack on metals is called **corrosion.** There is a considerable damage to metallic structure due to corrosion. *Corrosion of iron or steel is called* **rusting** *and takes place in presence of air and water vapour.* Rusting may cause collapse of bridges, leaking of roofs covered with steel sheets and may have other harmful consequences.

Iron and steel are the most widely used of all metals. Much work has been done on prevention of corrosion of these metals. Rusting can be prevented by covering the surface of iron and steel so that it cannot come in contact with air and water vapour.

A long lasting protective covering is a layer of zinc. The process by which zinc is coated is called **galvanising**. The sheet to be protected is cleaned with a dilute acid (this process is called pickling) and dipped into a bath of molten zinc. A thin layer of zinc adheres to the surface of the iron sheet. The layer of zinc may also be deposited by rolling iron in zinc dust at a high temperature. Zinc being more reactive than iron forms a coat of zinc oxide on the surface of iron and protects iron from further attack. Zinc oxide is also not porous to air and water. Even scratched galvanised iron will not corrode. However galvanised iron cannot be used for making food containers because the juices in food are acidic and dissolve zinc forming poisonous compounds.

TABLE 7.1. Composition, properties and uses of common alloys

Name of the alloy	*Composition*	*Use*	*Properties*
1. Brass	Copper, zinc	Cartridge cases, musical instruments, machine parts.	Easily moulded, resists corrosion.
2. Bronze	Copper, tin	Statues, coins, medals	Hard, can be easily cast.
3. Duralumin	Aluminium, copper, magnesium and manganese	Aircraft	Light and tough.
4. Magnalium	Aluminium, magnesium	Scientific instruments, aircraft	Light, tough and resists corrosion.
5. German silver	Copper, zinc, nickel	Utensils and decorative pieces	Silvery appearance and resists corrosion.
6. Solder	Lead, tin	Soldering joints, fuse	Low melting point, alloys with soldered surfaces.
7. Type metal	Lead, tin, antimony	Type used in printing	Expands on solidification.
8. Stainless steel	Iron, chromium, nickel	Cutlery and machine parts.	Resists corrosion.
9. Pewter	Tin, copper, bismuth and antimony	Carburettor, table ware	Casts easily and hard.

For food containers, steel is covered with tin. The process is called **tin-plating.** A tin plated container is protected from corrosion as long as the entire tinned surface is intact. A scratch on the tin plated can will cause the underlying steel to rust because steel is more reactive than tin.

Steel can be protected from rusting by **electroplating** it with nickel or chromium as is done in bicycle handles and car bumpers.

The most usual way of protecting iron girders and bridges from rusting is **painting**. Paints do not last long and hence repainting has to be done quite frequently.

If rust is to be removed from steel, phosphoric acid is often used.

Copper when exposed to moist air forms a green solid deposit on its surface. This deposit contains copper (II) hydroxide and is poisonous.

Lead is fairly resistant to corrosion and is used for making water pipes and coverings for electric cables. On account of its high density lead is used as protection against radiations and X-rays.

EXERCISES

1. (a) Which was the earliest alloy known to mankind? Give its composition. ______

(b) What are non-ferrous metals?

(c) Describe the principle involved in extraction of sodium metal.

2. (a) The metal zinc is extracted by roasting its sulphide ore followed by reduction. Write equations for the reactions taking place.

(b) Iron is extracted by putting coke and its oxide ore in a blast furnace. Write equations for the reactions taking place in the furnace.

3. (a) How can calcium metal be extracted?

(b) Why is aluminium much cheaper today than it was hundred years ago?

4. Give two uses of the metals mentioned below. Give reasons for each use.

(a) Copper

(b) Aluminium

(c) Lead

(d) Zinc

5. (a) What is an alloy?

(b) Give the composition and uses of the following alloys, mentioning the reason for each use.

(i) Brass

(ii) Duralumin

(iii) Magnalium

(iv) Solder

6. (a) Give two reasons each to show that magnesium and copper are metals. One of the reasons given must be based on a chemical property.

(b) What would you observe when iron nails are immersed in copper sulphate solution? Write the equation for the reaction.

(c) Mention a metallic oxide which can be reduced by heating. Write the equation for the reaction.

. (a) Using water how will you show that sodium, iron and copper have different reactivity. Write equations wherever possible.

(b) Both zinc and tin are used for checking the rusting of steel. When is zinc more suitable and when is tin considered to be better for checking the rusting of steel? ____

(c) How is a bicycle handle made of steel prevented from rusting? ____

8. Explain the terms:

(a) Malleability ____

(b) Mineral ____

(c) Ore ____

(d) Roasting of ore ____

8

Other Materials

8.1 Plastics

In the last fifty years many man-made materials have come to replace wood, metals, cotton wool and silk. Many of these are even used for making clothes. Look around your homes and you will see many articles made of substances called **plastics**, e.g. dishes, electric switches, curtains, bottles, cashmilon sweaters and clothes made of nylon and terylene.

The story of plastics is mainly that of 'giant molecules' called **polymers.** Some of these large molecules are found in nature and are modified to suit particular needs. Others are built synthetically by man. *Polymers are large molecules formed by linking together smaller, often identical, molecular units.* The process by which polymers are made is called **polymerisation** and the smaller molecular units which link together to form large molecules are called **monomers.**

Plastics are a group of substances which can be softened, and hence can be easily moulded,

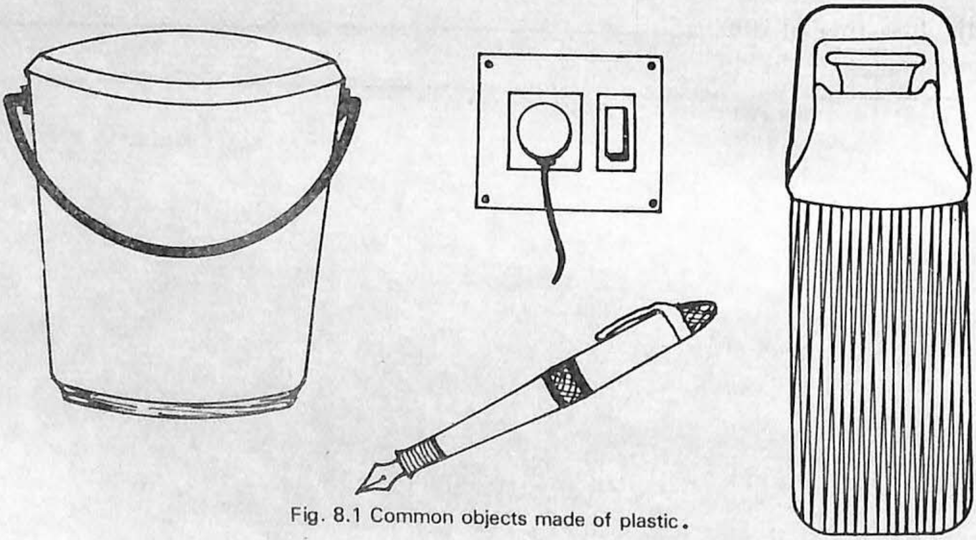


Fig. 8.1 Common objects made of plastic.

by applying heat and pressure. There are two types of plastics.

1. Thermoplastics

With this type of plastics the process of softening on heating and hardening on cooling can be repeated almost indefinitely, provided the plastic is not heated so strongly that it decomposes. Examples are polyvinyl chloride and nylon.

2. Thermosetting Plastics

This type of plastics can be softened by heating and moulded by pressure only once. Thus softening and moulding are irreversible, e.g. bakelite.

Many plastics are resistant to chemical attacks by atmosphere or by corrosive substances. Some are widely used as insulating materials in electrical industry and others are used as synthetic fibres.

Synthetic fibres, e.g. nylon and terylene, are made from plastics which are specially processed to form threads.

8.2 Common Plastics

1. Polythene

Polythene was earlier made by polymerisation of the compound ethylene at 1000 atmospheric pressure and a temperature of 200°C. Using a suitable catalyst, polythene can now be made at a much lower pressure. The polythene produced by this method has a high density.

Polythene is a semirigid, waxy, translucent solid. It is quite resistant to chemical action. Polythene bags have replaced paper bags because they are tough. It is also used for piping acids and other corrosive substances and for making flexible bottles. Another plastic which belongs to the same group as polythene is **polypropylene.** It is more rigid and heat resistant and is used for making gloves for surgeons, fishing nets and brushes.

2. Polyvinyl Chloride (PVC)

This is the most versatile and cheap plastic. It is made by polymerisation of vinyl chloride. It is obtained as a white powder which can be processed to give products which range from soft rubbery materials to hard sheets. A large amount of PVC is used as insulating covering for electric wiring. It is used for making flexible sheets. PVC can be made in the form of a paste which can be spread on cloth to make the cloth look like leather. It is used to cover motor car seats and furniture. It is also used to make hand bags and coverings for suitcases. Gramophone records, refrigerator linings and pipes are now made from PVC.

3. Polystyrene

It is one of the cheapest plastics. It is the third most important plastic. Articles made from polystyrene give a tinkling metallic sound when tapped. It is hard and brittle. It can, however, be made more tough and less brittle by mixing another substance (called co-polymer) during polymerisation. Polystyrene is mainly used for making packing materials, refrigerator parts, dishes, toys and heat insulators.

4. Bakelite

It is made by condensation of molecules of phenol and formaldehyde when a brownish powder called PF powder is obtained (P and F are initial letters of phenol and formaldehyde respectively).

PF powder is subjected to heating and pressure and put in moulds to give the required shape. Bakelite is not inflammable and is a good electrical insulator. It is used for making switches, plugs and elctrical fittings.

5. Perspex

This is an acrylic plastic. It forms transparent sheets like glass. It is unbreakable and is used for windows in aircrafts, wind screens of cars, street light fittings and telephones. Perspex can be dyed in different colours.

6. Celluloid

Celluloid is produced by treating cellulose with nitric acid when cellulose nitrate is formed. Cellulose nitrate is treated with camphor giving the plastic celluloid. This is not a synthetic plastic and is highly inflammable. It was used for a long time for making camera and motion picture films. It has now been replaced by newer non-inflammable plastics. Celluloid can also be used for making combs and soap boxes.

8.3 Synthetic Fibres

Synthetic fibres are also plastics which have been specially processed to form threads. These man made fibres now account for a large part of textile industry. These fibres are strong, crease resistant and are not attacked by moth. Clothes made of synthetic fibres should be carefully ironed because fibres are damaged by high temperature. Clothes made of synthetic fibres are easy to wash and maintain. Garments made of only synthetic fibres, however, do not allow the air to pass through them and are not comfortable for summers. These fibres are now blended with cotton and clothes made of these are suitable for all climatic conditions. The man-made fibres are also used to improve the properties of wool which is often made moth proof and crease resistant. You must have heard of terycot and terywool cloth. Some of the common synthetic fibres are as follows.

1. Rayon

When cellulose, obtained from cotton or wood fibres, is dissolved in carbon disulphide and sodium hydroxide solution a syrupy material called *viscose* results. When treated with dilute sulphuric acid, viscose solidifies into fibres, which are silky, have smooth feel and are strong. The product is called **viscose rayon** and is used for making light weight sweaters, stockings, conveyor belts, car tyre cords and many other industrial products.

2. Nylon

The commercial production of nylon was announced in 1938. It is widely used for making garments, socks, ropes, cords for aircraft tyres, parachutes, bristles of brushes, electrical insulators, conveyor belts and many other articles.

3. Terylene

Terylene appeared some years after the discovery of nylon. It has a texture more like wool and is more popular for suits when blended with cotton fibres (when it is called **terycot**) and wool (when it is called **terywool**). It is not so elastic and is, therefore, unsuitable for stockings. It is less absorbent.

4. Orlon, Acrilon and Cashmilon

Orlon, acrilon and cashmilon are newer synthetic fibres. These fibres have the warmth, bulk and feel of wool. They are considered as substitutes for wool. They are also characterised by resistance to chemicals,moth, sunlight and other insects. They are washable and shrink-proof.

Rapid growth of plastics will continue for many years to come. New uses will be found for existing plastics and new plastics will be discovered. It may even surpass the production of steel.

8.4 Glass

Glass has played an important role in the development of civilization. Without it there might have been no electric lights, no thermometers, no bottles, no cameras and telescopes and no chemical apparatus. The first maker of glass is unknown. He must have lived more than five thousand years ago.

Manufacture of Ordinary or Soft Glass

Glass is amorphous and is not a single substance but a homogeneous mixture of silicates of certain metals e.g. sodium and calcium. Ordinary glass is called **soda glass** and is made by heating a mixture of sand (silicon dioxide, also called silica), soda ash (sodium carbonate) and limestone (calcium carbonate). The heating is done in a rotary furnace at about 1400°C. A clear liquid is formed which is rolled into sheets. Hot viscous glass is blown through a blow pipe into bulbs, and many other glassware which we use in our laboratories. This process is called **glass blowing**. Bottles are, however, entirely machine made using moulds. Articles made of glass are cooled slowly. This is called **annealing.** If cooled rapidly, they break into pieces with slight disturbance.

Other Types of Glass

Glass having specific properties is made by changing the ingredients which are heated in the furnace.

1. Hard Glass: It is also called *potash glass* and is made by heating sand, potassium carbonate and calcium carbonate. It is thus a mixture of potassium and calcium silicate. It can withstand high temperatures without melting and is more resistant to acids.

2. *Flint or optical glass:* If lead oxide is used in place of sodium and calcium carbonate, a sparkling glass results. It is used for making lenses and prisms.

3. *Pyrex glass:* It contains boron oxide and aluminium oxide in place of calcium oxide present in ordinary glass. This glass can be heated over a flame, and cooled quickly without cracking. It is used for making laboratory equipments and cooking utensils.

4. *Shatter-proof glass:* It is made by placing a sheet of plastic such as celluloid between sheets of glass plates. In case of breakage, broken pieces stick to the plastic and are prevented from flying about. Bullet proof glass consists of several alternate layers of plates of plastic and glass.

5. *Ground glass:* It is obtained by rubbing soft glass with sand when glass becomes translucent. It is used in lamp shades for giving uniform light, and in houses and hospitals to avoid inside visibility.

6. *Glass wool*: It consists of thin fibres of glass and is used for filtration and as an insulating material.

Coloured Glass

Glass is coloured by adding various metallic oxides to molten glass during its manufacture. Different shades of colours are produced by using different quantities. These metallic oxides produce their silicates which, being coloured, impart their own colour.

(a) Cobalt oxide gives blue colour.

(b) Chromium oxide gives green colour.

(c) Selenium oxide and zinc oxide give a highly transparent red glass used for railway signals.

(d) A ruby red glass is obtained by incorporating gold in the glass.

(e) Manganese dioxide gives violet colour.

EXERCISES

1. (a) What are plastics? ______________________________

____

____

(b) What types of compounds are plastics? ______________________

____

2. Define:

(a) Polymerisation ______________________

____

____

(b) Thermoplastics ______________________

____

____

(c) Thermosetting plastics ______________________

____

____

3. (a) How are synthetic fibres superior to cotton fibres? ______________________

____

____

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(b) In which respect is nylon not suitable for making garments? ______________________

____

____

(c) What is terywool? ______________________

____

4. (a) How is polythene made? ______________________

____

____

____

(b) Give two uses of polythene. ______________________

(c) Name the plastics used for making gloves for surgeons.

5 (a) What is PVC?

(b) Mention three uses of PVC.

6. Which plastic material would you choose for making the following articles.

(a) Outdoor water pipes.

(b) A flexible bottle.

(c) Gramophone records.

(d) Packing material.

(e) A cloth which is a substitute for leather.

(f) A bottle for concentrated acids.

(g) Windows.

(h) As a substitute for wool.

(i) Cinema film.

(j) Switches.

7. (a) What is common glass?

(b) Is glass a compound with definite composition?

(c) Is glass crystalline or an amorphous substance?

8. (a) What is hard glass?

(b) How is it better than common glass?

9. (a) How is common glass manufactured?

(b) Why should articles made of glass be cooled slowly?

(c) What is glass blowing?

10. (a) What is pyrex glass?

(b) In what respect is it better than other glasses?

11. (a) How is ground glass made?

(b) Give one use of ground glass. ______________________

____

12. (a) How is bullet proof glass made? ______________________

____

____

(b) How is coloured glass obtained? ______________________

____

____

13. Name the compounds which will impart the following colour to glass:

(a) Red ______________________

(b) Green ______________________

(c) Dark Blue ______________________

9

Chemistry in Agriculture

During the twentieth century agriculture has made great strides. But to keep pace with the increasing population there is urgent need to increase food production still more. To achieve this aim, we must have knowledge of the condition of the soil and the nutrients which plants need. It is also essential to protect plants from pests and finally the grains should be properly stored in a dry place.

9.1 Soil

Soil plays the most important part in the growth of plants. It retains moisture and supplies food to plants in the form of mineral salts by dissolving them in water. It also permits the air to reach the roots.

For proper growth of plants soil should be neither too acidic nor too alkaline. A neutral soil has a pH of 7. The sample of soil to be tested is taken in a beaker. Some water is added and the contents stirred with the help of a glass rod. On filtering, the soluble acidic or alkaline substance present in the soil is collected as filtrate. The filtrate is tested with litmus paper or pH paper. Acidic substances turn blue litmus red and alkaline substances turn red litmus blue. The pH paper gives the actual pH of the solution. When the pH is less than 7, the soil contains acidic substances and a pH of more than 7 indicates alkaline soil.

Heavy rains help in turning soil acidic. To reduce acidity, the soil may be carefully treated with lime. Alkaline soil called *usar soil* is found in areas with low rainfall. Gypsum, a mineral containing calcium sulphate is added to reclaim usar soil.

Different plants need different optimum pH values of soil for their good growth. For a better yield of citrus fruits the soil should be slightly alkaline.

9.2 Nutrients for Plants

Manures and fertilisers are substances which contain mineral salts which are used by plants in manufacturing their food. When man began to harvest plants the mineral salts present in the soil were used up. To replace these, manures and artificial fertilisers had to be added to keep up and increase the yield of crops.

9.3 Manures

Manures are natural fertilisers obtained from dead plants and animal remains. There are two types of manures.

1. Green Manure

These are obtained by cutting down plants and burying them into the soil to decay. They enrich the soil with all the mineral salts which they have absorbed from the soil. For example, the plants of the bean family when mixed in the soil decay to supply nitrates to the soil. When plants die and decay they add to the soil a layer of organic matter called **humus,** which supplies food to new plants.

2. Animal Manure

Cow dung turned into compost, and powdered bones are rich in mineral salts.

These manures are called **organic fertilisers.** Although they are complete fertilisers, the nutrients present in them are not sufficient for plants.

9.4 Artificial Fertilisers

These are chemicals prepared in the laboratory and contain elements which are necessary for the growth of plants. The three principal elements taken by the plants are nitrogen, phosphorus and potassium. Artificial fertilisers are compounds containing one, two or more of these elements along with oxygen.

1. Nitrogen Fertilisers

Deficiency of nitrogen in the soil causes reduction in yield. The leaves change in colour from green to yellow. Several nitrogenous fertilisers are available. **Urea,** an organic matter also found in animal manures (such as urine of animals) is an excellent nitrogenous fertiliser containing the highest percentage of nitrogen. **Ammonium sulphate** produced by reacting ammonia with sulphuric acid is also used as nitrogenous fertiliser. **Ammonium nitrate** has a higher percentage of nitrogen than ammonium sulphate. It is a hygroscopic substance and hence absorbs water vapour from the atmosphere. It must be stored under dry conditions. Other nitrogenous fertilisers are **sodium nitrate** (chile salt petre) and **calcium nitrate.**

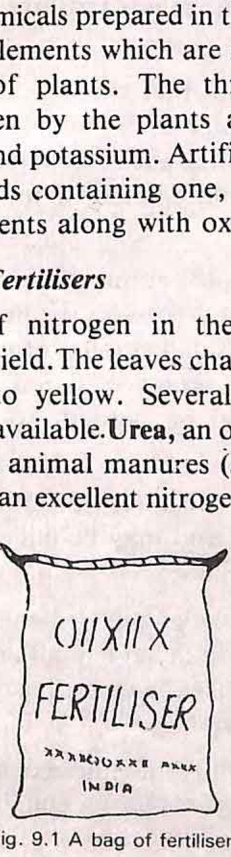

Fig. 9.1 A bag of fertiliser

$$\underset{\text{ammonia}}{2NH_3} + \underset{\text{sulphuric acid}}{H_2SO_4} - \underset{\text{ammonium sulphate}}{(NH_4)_2SO_4}$$

$$\underset{\text{ammonium hydroxide}}{NH_4OH} + \underset{\text{nitric acid}}{HNO_3} \rightarrow \underset{\text{ammonium nitrate}}{NH_4NO_3} + H_2O$$

The practice of adding liquid ammonia to irrigation water is another method of adding nitrogen to the soil and is gaining wide applications in the West.

At present about 100 million tons of nitrogenous fertilisers are being added to the soil each year all over the world. The nitrogenous fertilisers could double or treble the yield of grain crops.

Whereas urea can be conveniently used as a source of nitrogen to the soil, the choice of other nitrogenous fertilisers depends upon the nature of the soil. For example, ammonium sulphate increases the acidity of the soil and sodium nitrate increases the alkalinity of the soil. Leguminous plants such as pea and gram have nitrogen fixing bacteria in the nodules of their roots that directly absorb nitrogen from the atmosphere. This is available to plants and improves crop production.

2. Phosphorus Fertilisers

Phosphorus is an important element for the development of roots and structure of plants. It makes the plants sturdy and hastens the ripening of fruits. It increases the sugar content in beetroot, the content of starch in potatoes and the yield of grain crops. Phosphorus helps the plants to resist diseases.

Phosphorus is usually added in the form of some acid phosphate because calcium phosphate is very slightly soluble in water. Superphosphate fertiliser is obtained by treating rock phosphate with concentrated sulphuric acid.

$$\underset{\text{rock phosphate}}{Ca_3(PO_4)_2} + 2H_2SO_4 \rightarrow \underbrace{Ca(H_2PO_4)_2 + 2CaSO_4}_{\text{superphosphate}}$$

If rock phosphate is treated with phosphoric acid, triple superphosphate is formed.

$$Ca_3(PO_4)_2 + 4H_3PO_4 \rightarrow \underset{\text{triple superphosphate}}{3Ca(H_2PO_4)_2}$$

This is not diluted with calcium sulphate, which is almost insoluble in water, and contains about three times as much phosphorus as is present in superphosphate, if the same weights of the two are taken. Plants can easily absorb the triple superphosphate.

3. Potassium Fertiliser

Potassium is also essential for the growth of plants. It is used in the early stages of the growth of plants and assists in formation of carbohydrates, fats and proteins. It makes the plants stronger and disease resistant. Potassium sulphate, potassium nitrate and potassium chloride are the chemicals which enrich soil with potassium. When wood is burnt, the wood ash contains about 18% of potassium carbonate and can be used as potash fertiliser.

9.5 Mixed Fertilisers

When a fertiliser contains compounds which provide two or three of the elements, nitrogen, phosphorus and potassium, it is called a **mixed fertiliser.**

A fertiliser which contains compounds containing all the three elements is called **NPK type** of fertiliser (N = nitrogen, P = phosphorus, K = Potassium) and is a complete fertiliser. This is prepared by proper blending of various nitrogen, phosphorus and potassium compounds.

When concentrated ammonium hydroxide solution is added to superphosphate, the process adds nitrogen to phosphorus fertiliser at a low cost. It also checks the rotting of bags of fertiliser due to presence of unreacted sulphuric acid present in superphosphate. This is called **NP type** fertiliser.

NK type of fertiliser is a mixture of nitrogen and potassium compounds.

Nitrophosk is an important NPK type of fertiliser. **CAN** or calcium ammonium nitrate is another important fertiliser.

A mixed fertiliser of NPK type is more effective than an individual fertiliser.

9.6 Minor Fertilisers

In addition to nitrogen, phosphorus and potassium, some minor fertiliser constituents are essential.

1. It is found that the deficiency of *zinc* causes the change of colour in the leaves of many plants and as such zinc sulphate in small doses must be added to the soil for a better yield of crops.

2. *Magnesium* atoms are present in chlorophyll and catalyse photosynthesis. Though a very small amount of magnesium is needed, but its presence is vital. Magnesium sulphate may be added to meet this requirement.

3. Iron is also essential for proper photosynthesis and may be added as ferrous sulphate in small doses.

4. Borax provides the element boron. It is added to the soil in very small amounts. It is found to increase the quantity of vitamins in fruits and vegetables.

5. Besides those mentioned above, small amounts of copper, calcium and manganese are

also minor nutrients for plants. Their use is limited to specific types of plants.

Minor fertilisers are also sprayed on plants.

9.7 Precautions in Using Fertilisers

There must be plentiful supply of water available when fertilisers are added to the soil. Excessive use of fertiliser may damage the crop, so that care has to be taken to add the right quantity of fertiliser according to the need of the soil and the crop.

9.8 Plant Protection

Plants of crops get diseased when they are attacked by virus, fungi, insects and bacteria. About 20 per cent of crop is lost in this way. Weeds are unwanted plants which hamper in the growth of plants. They absorb some of the minerals in the soil.

Weed Killers

Herbicides are a group of chemicals used for destroying weeds. Examples of such substances are 2,4-D (2, 4 dichloro phenoxyacetic acid) or 2,4,5-T (2,4,5 trichloro phenoxy acetic acid).

Pesticides and Insecticides

Pests are animals and plants (like weeds) which hamper the growth of plants. The chemicals used for destroying insects are called **insecticides.** These and other chemicals used for destroying mites, fungi and rats are called **pesticides.**

The most commonly used pesticides are DDT (Dichloro diphenyl trichloroethane), BHC (Benzene hexachloride) and Gammexane (Hexachloro benzene). They are used to kill insects. Other pesticides are aldrin, parathion and malathion.

All these pesticides are highly toxic and they destroy the useful insects also. The use of insecticides may have harmful effects and they should be used carefully. Parathion and malathion are however relatively safer. Due to constant use of pesticides the insects may develop resistance to these. Pesticides are sprayed in the form of dust and as suspensions in a liquid.

Rats destroy food and spread diseases. They are destroyed with zinc phosphide which is mixed with foodstuffs to be eaten by rats.

Fungicides

Due to growth of fungus, leaves of plants and fruits are destroyed. The chemicals used to check the growth of fungus are called **fungicides.**

Powdered sulphur; sulphur dioxide and Bordeaux mixture (a mixture of copper sulphate and milk of lime) are good fungicides.

The pesticides and fungicides are washed down in the soil and affect the next crop. If they are not completely removed they are eaten by animals such as cows. They may ulimately reach man and harm him. Hence a judicious use of these is essential.

EXERCISES

1. (a) What are fertilisers? ______________________________

____

____

(b) What is the difference between manure and an artificial fertiliser?

(c) What is compost?

2. 5 g each of two samples of soil are taken in test tubes and water is added to both. The water is decanted and the following tests are performed.

   (a) When a red litmus is put into the decanted water of sample A it turns blue. What does this show?

   (b) When a blue litmus is put into the decanted water of sample B it turns red. What does this indicate?

3. (a) What is humus?

   (b) What is usar soil?

   (c) How is usar soil reclaimed?

4. (a) What major elements are needed for the growth of plants?

   (b) Name three important minor nutrients and their usefulness to the plants.

(c) Why is borax added to the soil?

5. (a) What are mixed fertilisers?

(b) What are NPK and CAN fertilisers?

(c) State three precautions which a farmer should take while applying fertilisers to the soil.

6. (a) What do you understand by plant protection?

(b) What are pesticides?

(c) Name three commonly used pesticides.

(a) How are rats harmful to farmers?

(b) How can they be destroyed?

(a) What is a fungicide? ______________________________

(b) Name three fungicides. ______________________________

(c) What are the harmful effects of pesticides and fungicides? ________________

9. What are the functions of the following: ____________________

(a) Nitrogenous fertilisers ______________________________

(b) Phosphorus fertilisers ______________________________

(c) Potassium fertilisers ______________________________

10. (a) What elements are useful for better photosynthesis in plants? __________

(b) What are weeds? Name two herbicides. ____________________

11. How are Ammonium sulphate and Ammonium nitrate prepared? __________

12. (a) How is triple superphosphate superior to superphosphate? __________

(b) If the pH of a soil is 8, what would you do to correct it? __________

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